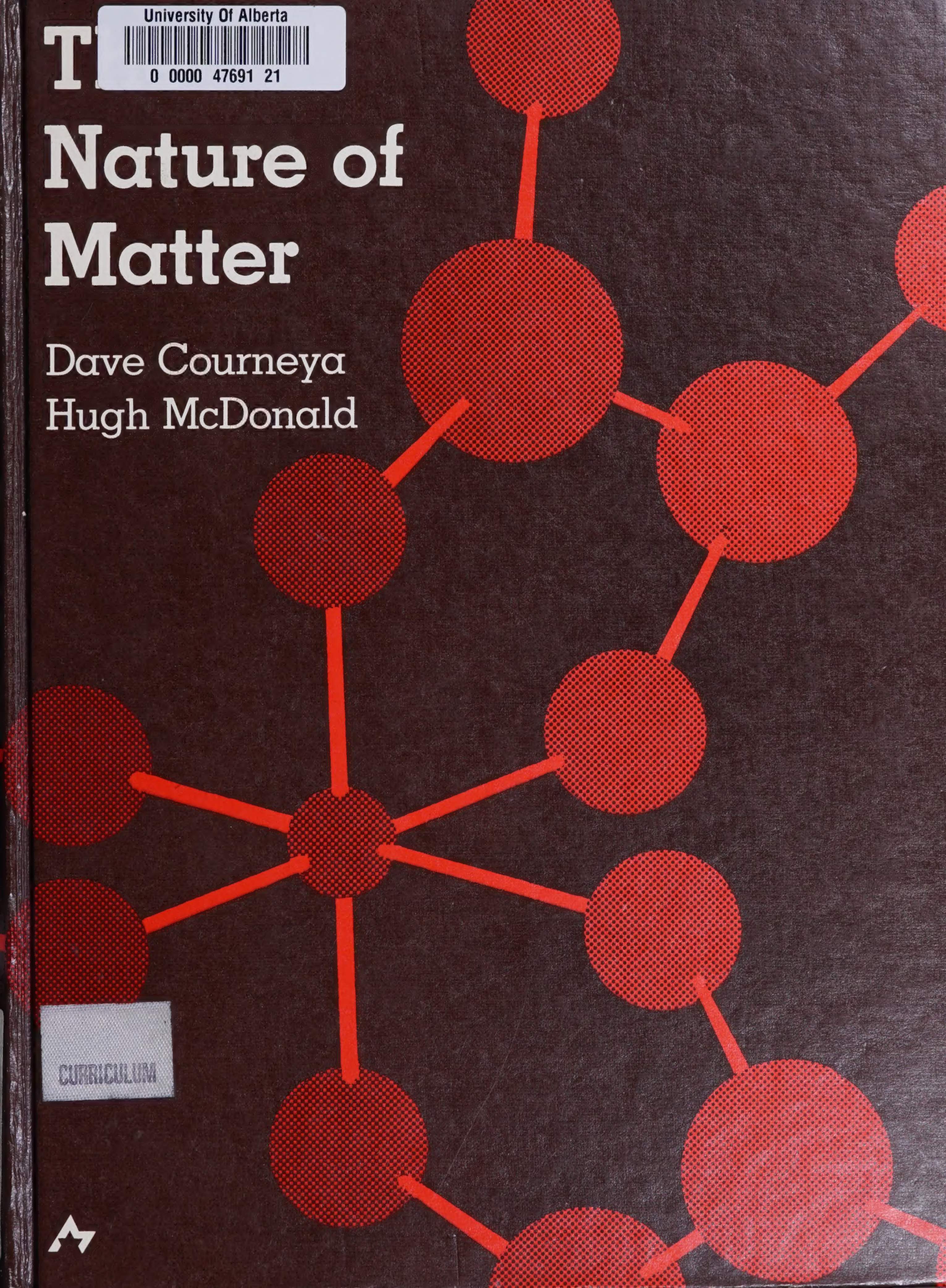
University Of Alberta
0 0000 47691 21

T Nature of Matter

Dave Courneya
Hugh McDonald



CURRICULUM

Heritage Fund

The purchase of this material was made possible through an Alberta Advanced Education and Manpower Library Development Grant from the Alberta Heritage Savings Trust Fund.

The Nature of Matter

Digitized by the Internet Archive
in 2021 with funding from
University of Alberta Libraries

https://archive.org/details/natureofmatter00cour

The Nature of Matter

Dave Courneya
Assistant Head of Science
East Northumberland Secondary School

Hugh McDonald
Head of Science
East Northumberland Secondary School

Canada Canada
metric métrique

D.C. Heath Canada Ltd. Toronto

This book is dedicated to the memory of Russ Haynes

The Metric Commission has granted the use of the National Symbol for Metric Conversion

Copyright © 1976 by D.C. Heath Canada Ltd.
All rights reserved. No part of this publication may be reproduced or transmitted in any form or by any means, electronic or mechanical, including photocopy, recording, or any information storage or retrieval system, without permission in writing from the publisher.
Printed in Canada
ISBN 0-669-00542-8

The authors would like to thank Steven Mills of D.C. Heath for all his help, and also Heather Sherrat for editing and Peter Maher for designing this book.

Bettman Archives, Inc., page 196
Brown Brothers, Sterling, Pennsylvania, page 159
Metropolitan Toronto Central Library, pages 143, 145, 153, 159, 197, 347
Miller Services, pages 17, 300
National Portrait Gallery, page 10
Other photographs: Peter Paterson
Illustrations: Peter Lole
Design: Peter Maher

LIBRARY
UNIVERSITY OF ALBERTA

D. C. HEATH, CANADA

ERRATA - *Nature of Matter*

UNIT 1

page 49 - The units of "C" - specific heat capacity are in joules per kilogram per degree kelvin - $\frac{J}{kg.\ K}$

page 68 - Replace solubility chart for gases with the following chart:

Solubility Given in Grams Dissolved in 100 g of water at a pressure of 101.3 kPa

	TEMPERATURE					
GAS	0°C	20°C	40°C	60°C	80°C	100°C
Carbon Dioxide	0.334 6	0.168 8	0.097 3	0.057 6	--	--
Hydrogen	0.000 2	0.000 16	0.000 13	0.000 11	0.000 07	0.000 00
Nitrogen	0.002 9	0.001 9	0.001 4	0.001 0	0.000 6	0.000 0
Oxygen	0.006 9	0.004 3	0.003 0	0.002 3	0.001 4	0.000 0
Nitric Oxide	0.009 8	0.006 2	0.004 4	0.003 2	0.001 9	0.000 0

UNIT 2

page 81 - table 2.4 last line - change "reagent to "reactant"

page 102 - line 7 should read $V = 100\ ml \times \frac{300K}{283K}$

page 127 - Exp. 2.14 line 1 should read "Use the results of experiment 2.4"

page 129 - lines 11,20,25 change "$\frac{1\text{-}kPa}{K\text{-}mol}$" to "$\frac{1.kPa}{k.mol}$"

page 131 - Question 2 - second sentence should read "When 5.45g of copper metal were made to react completely, 6.81g of cupric oxide were produced

UNIT 3

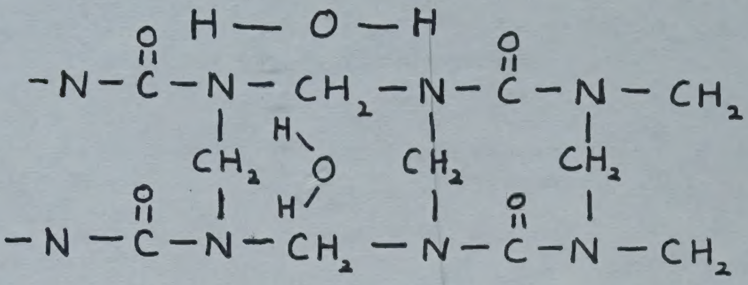
page 199 line 9 change ${}^{4}2\ He$ to ${}^{4}2\ He^{++}$
14 " He^{++} to $4\ He^{++}_{2}$
18 " H^{++} to $4\ He^{++}_{2}$

UNIT 5

page 273 question 5 - change "distillate" to filtrate"
304 reverse captions in diagram

UNIT 6

page 322 Warning - change "acetone" to "Acetylene" page 346 Note that the colour is missing and that the water molecules were not removed from bottom structure. The correct structure is:



UNIT 8

page 382 Lower left caption on diagram should read "Air Adjustment" instead of "Gas Adjustment"

UNIT 9

page 432 - Point 4 - Delete "as illustrated in the diagram"

Answers (Page 439-40)

page 48 - 1. 1 800 000 J page 68 - 19f 645g
page 105 - 4. (last two answers 5.09×10^5 ml; 0.056l
page 232 - 5. 125 ml page 256 - 5. 5×10^5 l
page 267 15. 70kg

Contents

1 An Introduction to Chemistry 18

Observations 20
Conclusions 23
Regularities 28
Conditions that Matter 31
Interpretations 32
Hypothesizing 34
Physical Change 37
States of Matter 39
Chemical Changes 40
Calorimetry 42
Solutions 49
Mechanical Mixtures 52
Properties of Matter 57
Colloidal State 59
Solubility 62
Solubility Curves 63
Questions and Problems 67
Some Words You Should Know 69

2 Matter and Gases 70

Phlogiston Theory 72
Changes in Matter 73
Laws Governing Reactions 75
Measurement and Chemical Change 83
The Nature of Matter 85
Gas Pressure 90
Gases and Temperature Change 101
Standard Temperature and Pressure 103
Gas Laws and Particle Model 106
Properties of Gases 107
The Elements 122
The Mole 123
Atomic and Molecular Masses 125
Molar Volume 127
Ideal Gas Law 128
Summary 129
Questions and Problems 131
Some Words You Should Know 133

3 The Atomic Model of Matter 134

The Molecular Model 137
Electricity and Matter 142
The Thomson Atom 144
Sub-Atomic Particles 149
Other Discoveries 151
The Rutherford Atom 152
The Bohr Atom 160
Law of Atomic Number 166
Electron Configurations 170
Conductivity of Solutions 176
Ions 179
Valence Numbers 182
Binary Compounds 186
Equations 186
Types of Bonding 191
Lewis Diagrams 193
Isotopes 195
Radioactivity 197
Summary 201
Questions and Problems 202
Some Words You Should Know 203
Some Names You Should Recognize 204

4 Descriptive Chemistry and Stoichiometry 206

Gas Production 209
Peroxides 211
Catalysis 212
Percentage Composition of Oxygen in Air 216
Acids and Bases 219
Naming Acids and Bases 222
Acid Strength and Base Strength 225
Neutralization 227
Molarity and Molar Solutions 229
Titrations 233
Displacement Reactions 242
Oxidation and Reduction 251
Carbon Dioxide 252
Gas Volumes from Equations 254
Stoichiometry 258
Summary 264
Questions 266
Problems 267
Some Words You Should Know 269

5 Man's Greatest Chemical Achievement 270

The Active Metals 272
The Halogens 275
Building the Periodic Table 289
The Inert Gases 291
Across the Periodic Table 293
Graphical Analysis 295
The Periodic Table 303
Questions and Problems 306
Some Words You Should Know 307

6 Organic Chemistry 308

Carbon 310
The Carbon Atom 315
Compounds of Carbon with Hydrogen 318
Branched Molecules 324
Functional Groups and Chemical Behaviour 329
Aldehydes and Ketones 334
Carboxylic Acids 336
Esters 339
Naming Organic Compounds 342
Plastics 343
Summary 350
Questions and Problems 351
Some Words You Should Know 353

7 The Marketplace — Applications of Chemistry 354

Of Consumer Interest 356
Hardness in Water 356
Soaps, Detergents, Bleaches and Fragrances 360
Photography 365
Properties of Metals 368
Qualitative Analysis 371

8 Laboratory Skills 380

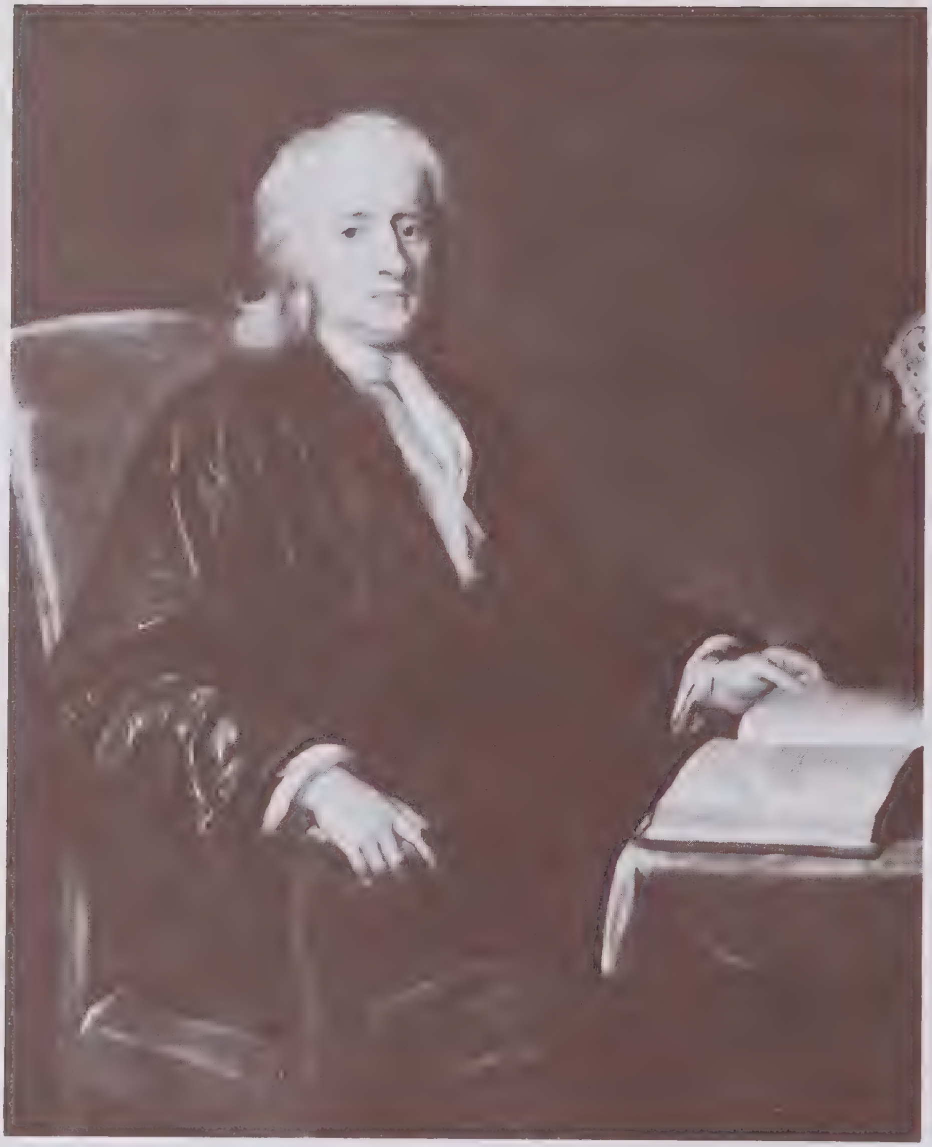
Using the Bunsen Burner 382
Working with the Balance 384
Cleaning Glassware 387

Filtration 387
Measuring Volumes 389
Glass Tubing Manipulations 392
Handling Liquids 395
Identifying Odours 397
Melting Point Determination Using a Capillary Tube 398
Heating in a Crucible 399

9 Independent Study Projects 400

Introduction 402
Drugs 403
Food Chemistry 405
Nutrients in Milk 407
Molecular Masses 409
Industrial Processes 411
Cosmetic Chemistry 412
Measuring Molecular Size 413
Measuring Inter-Molecular Energies 415
Measuring the Strength of a Bond 416
Pulp and Paper 417
Indicators 419
Automobile Chemistry 420
Crystals 422
Chromatography 424
Water Pollution 425
Air Pollution 427
Smoking 428
Dyes and Dyeing 430
Electroplating 431
Electrolysis 433
Other Projects 435
Answers 439
Index 441

The
Nature of
Matter



Newton was born in 1642, the same year that Galileo died, and died in 1727. He was a giant among scientists, having left us a legacy that includes calculus, mechanics, the law of universal gravitation and optics.

Perhaps his greatest gift was to refocus the attention of science on the need for experiments to supply the data necessary for knowledge to grow on.

Newton's statue at Trinity College, Cambridge is inscribed (in Latin) —

"He who excelled the human race in the power of thought."

To the Student

Isaac Newton is quoted as saying, "I do not know what I may appear to the world; but to myself I seem to have been only like a boy, playing on the seashore, and diverting myself in, now and then, finding a smoother pebble or a prettier shell than the ordinary while the great ocean of truth lay all undiscovered before me."

In one sense at least we are all Isaac Newtons. We are all looking for the undiscovered truths around us. We believe that the only way to find out is by looking. You must peek and poke and prod as you try to unearth new information that will help you to understand the world better.

Newton also said, "If I have seen further it is by standing on the shoulders of giants."

In this sense too we are all Newtons because we are all able to stand on the shoulders of the Einsteins, the Newtons and the Daltons who have gone before.

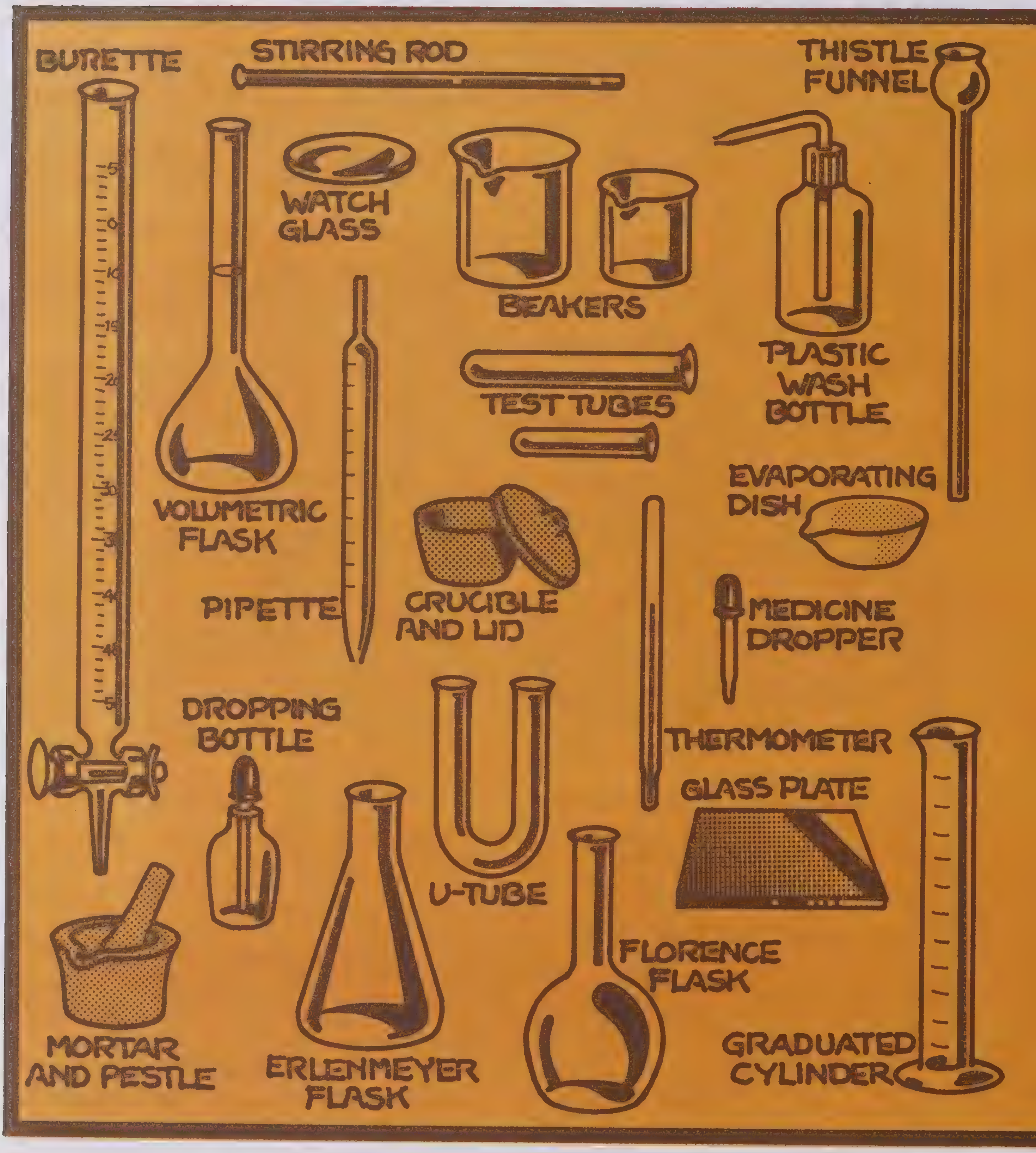
This course in chemistry has been designed so that you may learn not only some of the facts of science but also how to use those facts. For many of you, this latter skill will be the more important.

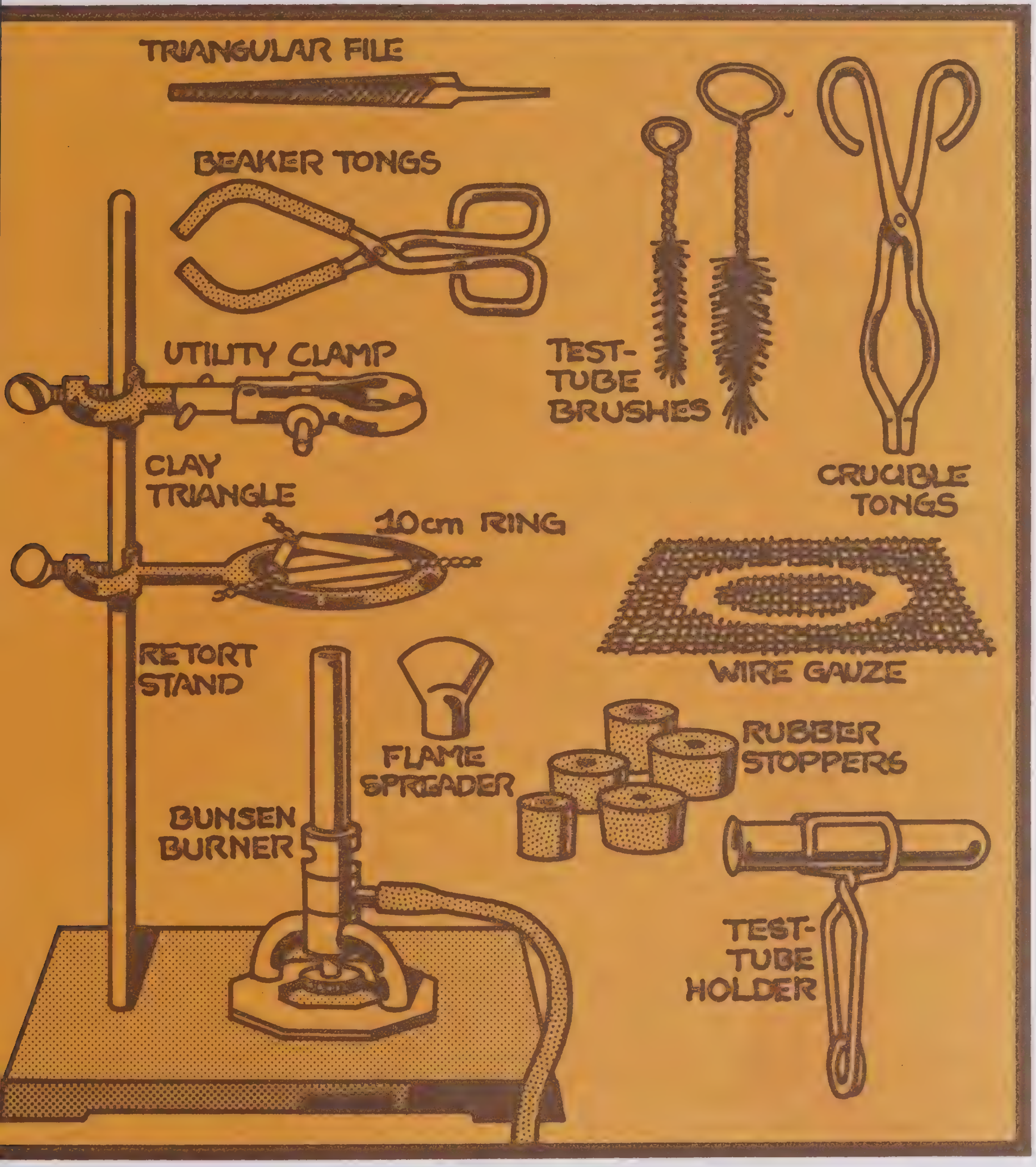
We hope that along the way you will learn to enjoy the ghostly presence of the Newtons, the Daltons and all the others as they peer over your shoulder into your test tubes. Try to sense the presence of Lavoisier's headless body or Madame Curie's pale and drawn face. Listen for Mendeleev as he mutters in Russian or Berzelius as he jokes in Swedish.

Make the course enjoyable by knowing in advance what you are going to do in the laboratory. Prepare a data sheet to save yourself some valuable laboratory time. Record each event as it occurs and in as much detail as possible. Try to follow the directions as accurately as possible. Lastly, try to be conscious of the fact that all chemicals can be dangerous if they are not handled with care.

Know where the fire extinguisher is and how to use it.
Know where the first-aid kit is kept.
Know where the eye bath is and how to use it.
Keep long hair tied back and away from burner flames.
Clean up chemical spills immediately.
Clean up broken glass immediately.
Always put the tops back on bottles.
Always follow instructions, both written and verbal, carefully.
Remember that things that have been held in a burner flame get hot ... and stay hot for a long time!
Learn to identify the equipment shown on the next pages. In the event that a chemical is spilled on your hands, be sure to wash it off immediately with lots of cold water.
Be sure never to try anything your teacher has not approved of in advance.

Common Laboratory Equipment


BURETTE
STIRRING ROD
THISTLE FUNNEL
WATCH GLASS
BEAKERS
PLASTIC WASH BOTTLE
TEST TUBES
VOLUMETRIC FLASK
EVAPORATING DISH
PIPETTE
CRUCIBLE AND LID
MEDICINE DROPPER
DROPPING BOTTLE
THERMOMETER
GLASS PLATE
U-TUBE
FLORENCE FLASK
GRADUATED CYLINDER
MORTAR AND PESTLE
ERLENMEYER FLASK


TRIANGULAR FILE
BEAKER TONGS
UTILITY CLAMP
CLAY TRIANGLE
10cm RING
RETORT STAND
BUNSEN BURNER
FLAME SPREADER
TEST-TUBE BRUSHES
CRUCIBLE TONGS
WIRE GAUZE
RUBBER STOPPERS
TEST-TUBE HOLDER

Foreword to the Teacher

Ask any ten people who are reasonably well versed in science to name the top ten scientists of all time, then compare the lists of names.

At least one name will be common to all the lists—Isaac Newton. But Newton was a physicist and mathematician. Why bring up his name in connection with chemistry?

There is one compelling reason for all of us to be grateful to Newton. He himself is quoted as saying, "There is no arguing against facts and experiments." (E.N. da C. Andrade in *Sir Isaac Newton–His Life and Work.)*

Newton seems to be saying that experiments produce facts and facts are *not* arguable. There may be several different opinions on the interpretation of what the facts mean but we may not argue against a fact. Sometimes of course statements that are presented as facts are shown to be false, in which case they were not facts to begin with.

Many of the scientists in Newton's day believed that the best way to find the solution to a problem was to quote Aristotle or Socrates or some other respected figure from the past.

When the Royal Society of London, England was formed in 1662, it took as its motto: "Nullius in Verba," which translates freely as "We don't take anybody's word for it." (E.N. da C. Andrade.)

What they were trying to say was, "Don't talk about it, show us the results of your experiments!" They were trying to emphasize the need for experiments to produce results that could then lead to a conclusion or interpretation.

In Newton's day the old idea of quoting someone else was slowly yielding to the new idea of gathering facts and then trying to interpret them.

In essence, that is what this book is all about. We don't ask students to believe us. We want them *first* to look at the data and *then* to see how it might be interpreted. We, of course, have the benefit of many wise men of the past to help us in our deliberations, but the point remains the same. We are trying to interpret the facts.

We hope you will teach *science* from this book. That may sound trivial until you recognize that as far as we are concerned the process of gathering and interpreting the facts is just as important in a science course as the facts themselves. Perhaps the one thing that all the great scientists had in common was the ability to look at the facts and then interpret them. That is what science is all about. That is what we must teach young people to do.

Throughout this book we will use SI (Système International d'Unités) terminology and symbolism. For example, in SI the accepted unit of pressure is the pascal (Pa): 1 atm = 101 kPa.

There still seems to be some confusion over the SI units for

The Royal Society

concentration. SI calls for concentrations in chemistry to be expressed in moles per cubic metre (mol/m^3), but molarity (mol/ℓ) will continue to be an acceptable unit. We will therefore use molarity (M) as our unit of concentration.

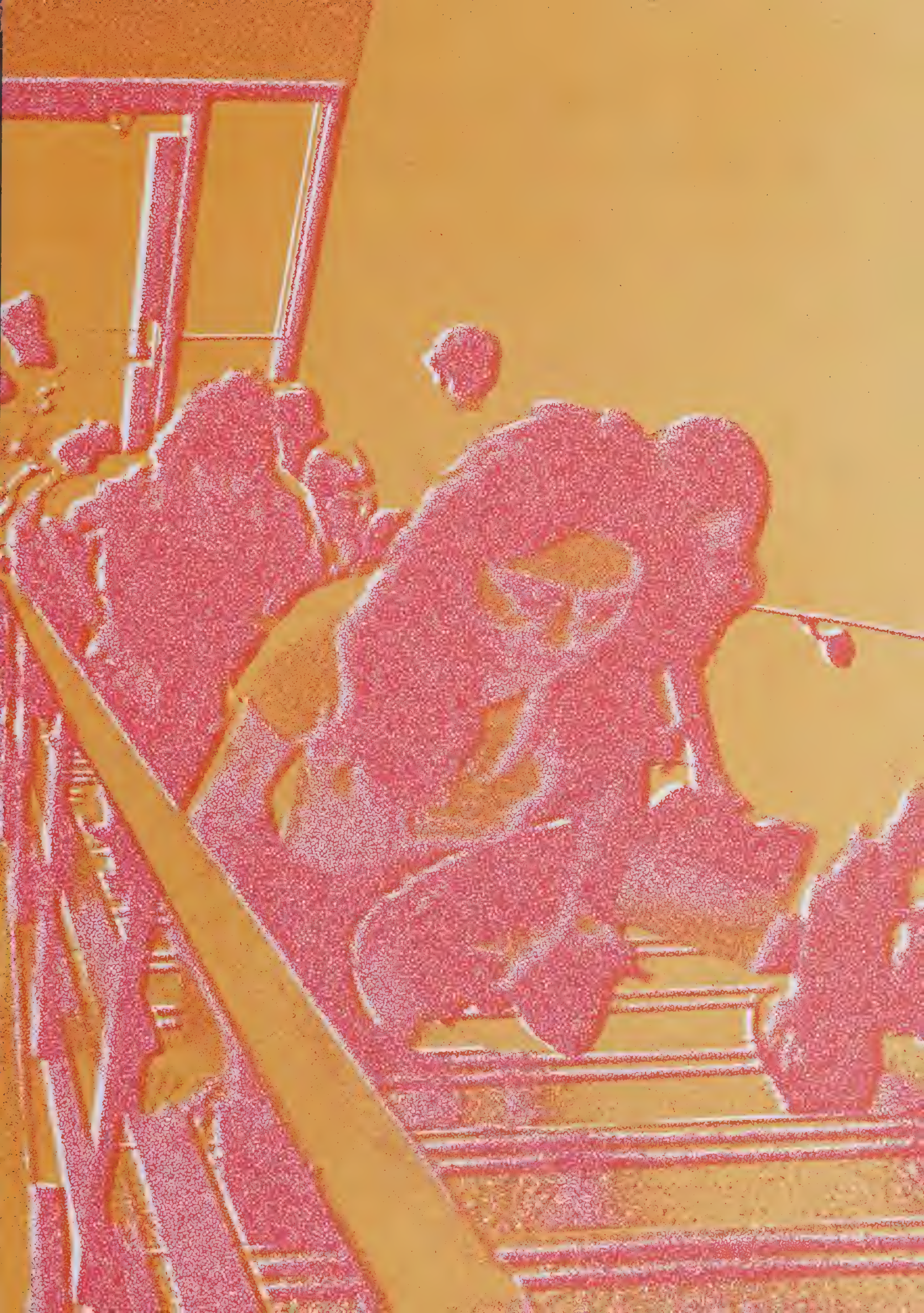
In all other areas we have tried to make our conversion to SI complete. The only exception will be found in Experiment 2.4 where we will continue to use millimetres of mercury as our unit of pressure for what we hope are obvious reasons. We feel that using pascals in this experiment would only cloud the issue and would result in unnecessary confusion for the student.

Galileo Galilei's work with telescopes helped him to verify the Copernican idea that the sun and not the earth was the centre of the universe. He is little remembered for his part in the revival of atomism. His proposals on the mass, shape and velocity of atoms fit in well with the modern model but were not widely accepted in his own day.

Leucippus, Democritus, Epicurus, Archimedes and many other ancient philosophers spent much time *thinking* about matter. Few actually did any experiements or gathered evidence to support their notion that matter was composed of small particles. These atomists were overshadowed by Aristotle and others who believed that matter was composed of earth, air, fire and water. It was not until the 1600s that atomism surfaced again as a reasonable model for matter.

1

Observations/
Regularities/
Conditions that
Matter/Generalizations/
Physical Change/
Chemical Change/
States of Matter/
Properties of Matter/
Mechanical Mixtures/
Solutions



Observations

1.1

Candles

This first experiment uses a very simple common object as the principal piece of equipment. You will gather some data, organize your data, evaluate your data and finally try to decide if it means anything. That, in essence, is what science is all about; trying to find out the meaning of things by looking for facts. Let us try and see what we can find out.

candle
matches
tin lid
balance
ruler

1. Describe an ordinary paraffin candle.
2. Fix the candle to a tin lid by softening the end of the candle in a flame and pressing it on the lid.
3. Measure the height of the candle and lid very carefully. Record the height.

4. Determine the mass of the unlit candle and lid carefully. Record the mass.
5. Light the candle. Allow it to burn for ten full minutes. Record as many new observations as you can during this burning period.
6. Carefully extinguish the candle. Allow the liquid material to cool.
7. Determine the mass of the candle and lid once more. Remeasure the height. Record these values.
8. Prepare a list of *changes* that took place during this experiment.

Questions

1. How many of your senses were you able to use in this experiment? How many senses do you have?
2. Does it make any difference to your observations how you hold the candle:
   (a) before you light it?
   (b) after you light it?
3. Can you explain the changes that took place during the burning process?
4. What conclusion can you draw from this experiment?
5. Are there any questions you do not have an answer for?
6. Where, in this experiment, did you:
   (a) gather data?
   (b) organize data?
   (c) evaluate data?
   (d) interpret data?

Almost automatically you should see the next step a scientist should take. He would try to find the answers to those questions that this experiment did not answer.

About 300 years ago, a man named Georg Ernst Stahl performed an experiment similar to the one you have just completed. Like you, Stahl and many other scientists of the day observed that the candle got shorter as it burned and that it also lost mass.

Stahl attempted to explain these two facts in the following way. He suggested that candles gave off something as they burned — something that had mass. He called this substance *phlogiston*.

Further experiments showed that wood, paper and many other substances lost mass when they burned. Stahl expanded his theory to suggest that any substance that burned gave off phlogiston.

It seemed only a small step to suggest that if a substance gave off phlogiston as it burned then it must contain phlogiston as part of its structure. Stahl was then able to speculate that since all substances that burn gave off phlogiston it must be the phlogiston that made them burnable!

The essence of Stahl's theory was simple:

1. Substances that burn contain phlogiston.
2. Substances that do not burn do not contain phlogiston.
3. Phlogiston must be present in a substance before it can burn.
4. Phlogiston is given off as a substance burns.
5. Phlogiston has mass.

1.2 More Candles

The next experiment is simple to perform. You will continue your investigation of a candle in an effort to understand more about burning (and perhaps about phlogiston).

2 candles — 1 long and 1 short
matches
1000 ml beaker
600 ml beaker
400 ml beaker
any other sizes of beaker available
timer or wrist watch with second hand
asbestos mat

1. Stand the short candle on an asbestos mat or the steel base of a retort stand. Light the candle.
2. Cover the candle with the 1000 ml beaker. As the confined candle burns observe the flame and the walls of the beaker carefully. Record everything as accurately as possible.
3. Remove the beaker, relight the candle and replace the beaker. This time, record how long it takes for the candle to "do its thing."
4. Repeat the procedure several times using different sized beakers. Record the time in each case in a table like Table 1.1.
5. Repeat the procedure again using the long candle instead of the short one. Record the time in the appropriate place in the table.

Table 1.1

	Time in Seconds	
Beaker Size (ml)	*Short Candle*	*Long Candle*
1000 ml		
600 ml		
400 ml		
etc.		

Questions

1. What ultimately happens to a burning candle that is covered with a beaker?
2. How is this related to the size of the confining beaker? Show this on a graph.
3. Was there any evidence of a substance that could have been phlogiston?
4. Why does a candle go out when it is confined? Is there more than one possible explanation? How many different explanations can you think of?
5. What is the effect of the length of the candle on the times? Do you think it would have been better to use only one candle? Why?

Conclusions

We have just completed the initial step that any scientist must take in an effort to understand his world. We have been engaged in gathering data. The data that we have gathered is related to a candle. Now a candle is a very simple, ordinary object. What would a scientist find in a candle that could possibly be of interest? Let us consider for a moment one of the things that happened in the course of your experiment. When you covered the burning candle with a beaker, you observed that some time later the flame of the candle "died." Why would a candle be extinguished when covered with a beaker? How many different theories could explain this single fact? Let us consider a few explanations.

Explanation 1

The candle died of loneliness. Is this a reasonable explanation? We have all heard of animals dying of loneliness. We may even have heard of people who died of loneliness, but could the candle die of loneliness? How could we test this hypothesis? Perhaps the simplest test would be to place two or three or four or five candles under the beaker. If the candles continue to burn, then our explanation may have been correct. If the candles are extinguished, then presumably it was incorrect.

Explanation 2

The candle "starved" to death. Is this a reasonable explanation? This implies that the candle uses something inside the beaker, and when that is used up, the candle no longer burns. It needs something that is inside the beaker. Could it be something in the air, in the glass, or in the candle itself? How could we test this explanation? This time, you must design the test.

Explanation 3

The burning candle produces something that "chokes" it to death. Now this explanation implies that a burning candle produces something. It also implies that this is poisonous and when it builds up enough it "chokes" the burning candle. How could we test this explanation? How many other explanations are there? See if you can list two or three more explanations. Once you have given an explanation, you must suggest an experiment to test it.

In fairness, we should test all explanations, not just one or two. We should gather as much data as we can to try and determine why a candle goes out when it is covered with a beaker. Once we have gathered the data, we may have to organize it into a table or a graph. We may have to take some of the data and throw it out because the experiment did not produce the kind of data that we wanted. Once we have data that we consider reliable, we must then interpret it. That is, we must try and decide which explanation is the best. This exercise will illustrate the kind of thinking process that a scientist must go through when he attempts to solve a problem. The process can be summarized in the following way:

1. gathering data;
2. organizing data;
3. evaluating data;
4. interpreting data.

These are the kinds of things that scientists do in their efforts to solve the problems of the scientific world. As a student scientist, these are the things that you too must learn to do as you try to discover new ideas, new facts and new theories about the world you live in. Just one more point. As a scientist gathers more data and looks for new explanations, he usually finds new problems. In fact, the search for one answer usually produces two or three questions that have to be answered. As the scientist proceeds with his work, he normally generates two things:

1. the answers to some questions;
2. more questions.

Probably both of these are equally valuable. While it is satisfying to get answers to questions, it is perhaps more important to have more questions. The more questions there are today, the more answers there will be tomorrow. The more answers there are tomorrow, the better our understanding of the world will be.

1.3

Why does a Candle . . . ?

What follows next is a series of short simple experiments—maybe we could call them mini-experiments. Each of them will tell you something about candles and candle flames. Do these experiments in whatever order you choose. As you complete each part stop and ask yourself questions like: What does this tell me about the problem? Which idea does it seem to suggest is correct? Why does a candle flame go out when confined?

candle
tin lid
pneumatic trough
glass jar
beaker
2 250 ml erlenmeyer flasks
stoppers
glass droppers
rubber tubing
glass tubing
wood splints
cardboard squares

cobalt chloride paper
limewater (30 ml)
copper wire
aluminum foil

1. Fix a candle to a tin lid. Stand the candle in a trough of water. Light the candle. Lower a jar over the burning candle until the mouth of the jar is well below the surface of the water.
2. Cover a burning candle with a clean dry beaker. After the candle is out wipe the inside of the beaker with a piece of cobalt

chloride paper. Touch a second piece of cobalt chloride paper to a drop of water.

3. (a) Cover a burning candle with a 250 ml erlenmeyer flask. When the candle goes out, quickly turn the flask upright on the table and stopper it. Add 10-15 ml of limewater to the flask and shake.
   (b) Add 10-15 ml of limewater to a second flask filled with air and shake it.
   (c) Remove the stopper from the second flask and blow into it. Reseal and shake. Repeat several times if necessary.
4. Blow air from your lungs into a candle flame through the glass part of a dropper attached to a rubber tube.
5. Hold a beaker filled with cool water in the yellow part of the candle flame for a few moments.
6. Wind a copper wire around a pencil to make a coil about 2 cm long. Remove the coil from the pencil. Slowly lower the coil over the flame and wick.
7. Hold one end of a 8 cm long piece of glass tubing in the dark inner region of the flame; hold a lighted splint at the other end.
8. Blow out a candle that has been burning for a minute or so with a quick sharp puff. Try to relight the smoke.
9. (a) Extinguish the flame by pinching it between your thumb and forefinger.
   (b) Extinguish the flame by blowing it out. Try blowing at different velocities.
10. (a) Try to light a piece of wax that has no wick.
    (b) Try different materials as wicks; for example, paper match, sliver of wood, copper wire.
11. Take a 5 cm square of aluminum foil and cut a slot from one side to the centre. Slide the slot over the base of the wick of a burning candle.
12. Introduce a few small solid particles into the liquid in the bowl of the burning candle. Watch the particles carefully. (Use carbon from a burnt match.)
13. (a) Hold a 10 cm square of cardboard vertically in the flame for several seconds; remove it before it starts to burn.
    (b) Hold a square of cardboard horizontally in the flame just above the tip of the wick; remove it before it starts to burn.
14. Place a few drops of water at the base of the wick of a burning candle.

Questions

1. Does a candle produce something as it burns? What are these substances?
2. Does the candle consume something as it burns? What are these substances?

3. How many ways can you think of to extinguish a candle flame?
4. What basic principles are involved in extinguishing a candle flame?
5. What is required for the flame to continue burning?
6. Burning is often said to be a *dynamic* process. Can you explain the use of the word dynamic?

Problems

1. List the senses that could be used in making observations.
2. Some chemists say that the first law of chemistry is "Never lick the spoon." Can you explain why?

3. The following statements were made by a group of students examining an ordinary ball point pen:
   (a) It is an approximately cylindrical object.
   (b) The longest dimension is 14 cm.
   (c) The body of the pen is made of plastic except for the clip.
   (d) The body of the pen is black.
   (e) The circumference is approximately 3 cm.
   (f) The body of the pen can be taken apart in two pieces.
   (g) The top half of the pen is separated from the bottom half of the pen by a chromium plated iron washer.
   (h) The silver-coloured button on the top end of the pen may be depressed causing the writing tip to be retracted within the body of the pen.
   (i) Depressing the silver-coloured button produces a click.
   (j) The writing tip is a spherical steel ball.

   From this list of statements select:
   (a) the qualitative observations;
   (b) the quantitative observations;
   (c) the conclusions or interpretations.

4. Look carefully at the photograph of a classroom on page 27. Can you answer the following questions:
   (a) What country was the picture taken in? Can you be sure?
   (b) What subject is being taught?
   (c) What time of year is it?
   (d) What time of day is it?
   (e) Is the teacher a man or a woman?
   (f) What else can you deduce from the photograph?

Regularities

As you studied the observations made of a burning candle you probably found it very difficult, if not impossible, to draw any conclusions about the candle.

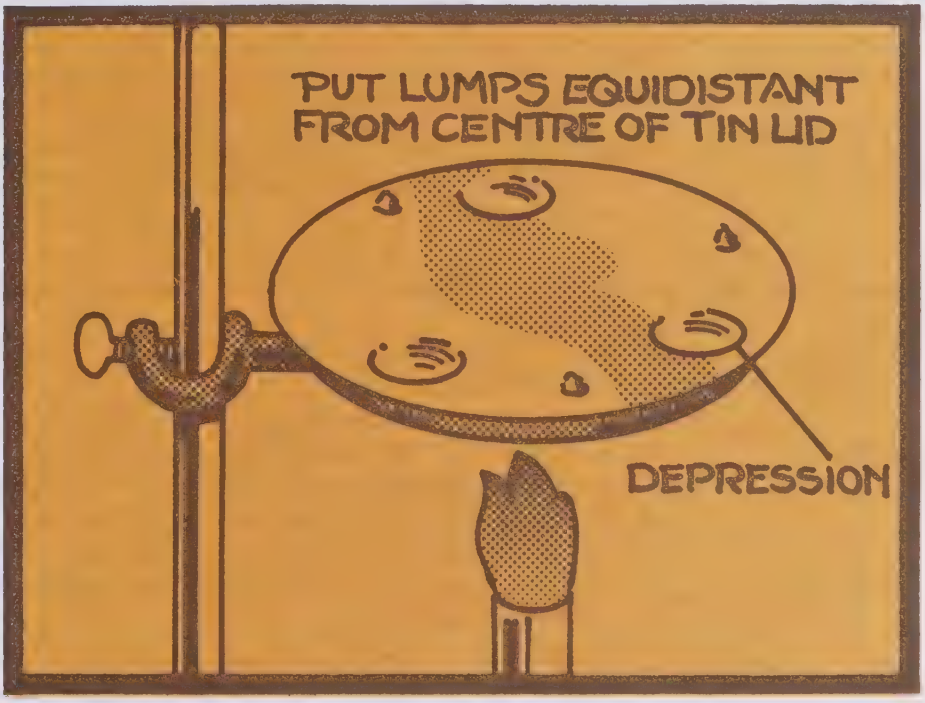
Why, even with a very detailed list of observations, should this be so? Is it possible that there is still something that you do not know about a burning candle?

Think of the questions that arose in your mind from Experiment 1.1. Perhaps questions come to mind like: What is the liquid in the bowl of the candle? Is the liquid in the bowl of the candle the same as the candle material?

These are questions with a ready answer, an answer that we will get by experiment.

1.4a Heating Solids

lid from a large juice can
retort stand and ring
wire gauze
candle
bunsen burner
250 ml beaker
candle material
sulphur
solder
lead
table salt
iron



1. Obtain a tin lid from a large juice can. The lid should have at least three depressions in it near the outer edge and all the same distance from the centre.
2. Support the lid on a retort ring with the ring about 20 cm from the base of the stand.
3. Obtain small lumps of the materials listed above; all lumps should be about the same size and shape. Place the lumps of material around the edge of the lid, the sulphur, candle material and solder being placed in the depressions.
4. Heat the lid from the centre with a candle flame. It is important that all materials be the same distance from the source of heat. Observe carefully and record all observations.
5. When there is no further change from the heat of the candle flame replace it with a bunsen burner. Heat the lid with a low flame at first; then increase the intensity of the flame. Heat with the hottest bunsen flame for at least 2 min.

Questions

1. List the substances in the order in which they began to melt. Do your classmates agree with you?
2. Do you think that every group should have the same order of melting? Why?
3. (a) Which substance(s) melted from the heat of the candle flame? Why did the other substances not melt?
   (b) Which substance(s) melted from the heat of the low bunsen flame? Why did the other substances not melt?
   (c) Which substance(s) melted from the heat of the high bunsen flame?
   (d) Why did some substances not melt at all in this experiment?
4. What conclusion(s) can you draw about the heating of substances?
5. What conclusion can you draw about the relationship between temperature and the solid state?
6. What is the temperature that appears in this experiment called?

1.4b

1. Put 50 ml of water into a 250 ml beaker. Support the beaker with a ring and wire gauze on your retort stand.
2. Place a lump of the solidified material from the bowl of the candle and a lump of unmelted candle material in the beaker onto the surface of the water. (The lumps should be the same size and the same shape.)
3. Gradually heat the beaker with your bunsen burner until the lumps just begin to melt. Watch closely now and note the order in which the two lumps begin to melt.

Questions

1. What was the order in which the candle material and the solidified liquid began to melt?
2. Based on your conclusions to Part A and the order of melting from Part B, what can you conclude about the candle material and the liquid from the bowl of the candle?
3. List the characteristics of:
   (a) the candle material;
   (b) the solidified liquid from the bowl of the candle.
4. How do these characteristics influence the decision that you made about the candle material and the solidified liquid from the bowl of the candle?

Conditions that Matter

A person observing a burning candle may have recorded as observations that the candle flame continually fluttered or that the candle material was a dull gray in colour. Perhaps the candle was observed near an open window with the lights out. A breeze from the window would explain the fluttery flame; the darkness of the room would explain the apparent colour of the candle material.

The conditions under which an experiment is performed are extremely important. Selecting improper conditions may result in misleading observations.

The most ideal conditions for an experiment should be selected, then carefully controlled.

In the example above the windows should have been closed and the light should have been turned on. These are clearly more ideal conditions.

Problems

1. When you observed the burning candle did you consider these conditions that mattered:
   Were the windows closed?
   Were you breathing on the flame?
   Were people walking by causing air currents?
   Were the lab fans on?
   Were the lights on?

These are all conditions that may have influenced the results of your experiment; these are all conditions that should have been controlled.

2. Think again of the candle experiment — were there any other conditions that should have been controlled?
3. Consider now the tin lid experiment. Did you think to control conditions that mattered?
   (a) For example, what would happen if the burner flame were not positioned under the centre of the lid? Is this a condition that mattered?
   (b) What would happen if the material on the lid were placed just anywhere, not around the outer edge? Is this a condition that mattered?
4. Here are some other conditions; pick out those that mattered in the tin lid experiment.
   (a) The windows were open.
   (b) The lights were on.

(c) People were walking about the room.
(d) The lab fans were on.
(e) The observer was breathing on the lid.
(f) The lid was only 2 cm from the bunsen burner barrel.
(g) The lumps of material were of different shapes.
(h) The depressions in the lid were not all the same depth.
(i) The lumps of material were of different sizes.
(j) The beaker was three-quarters full of water.
(k) The atmospheric pressure was 100 kPa.
(l) The air supply of the burner used was controlled by a lever at the bottom.
(m) The lid was not from a juice can.
(n) The room temperature was 24°C.
(o) The order in which the materials were placed on the lid was not the same for each lab group.

Interpretations

1. Not long ago the great forensic scientist, Earl N. Meyer, went to a seaside resort for his holidays. He was employed by the Metropolitan Police Department and this was his first holiday in several years. He was looking forward to a rest and was determined to enjoy himself even though it was still early in the season.
2. Lounging beside the pool one day, his eyes half-closed and the warm spring sun beating down, he was suddenly brought back to reality by a loud piercing scream from somewhere nearby. Springing to his feet he dashed toward the source of the sound, all thoughts of his holiday gone from his mind. He was not the first to arrive at the scene. Two ashen-faced young ladies, both dressed in identical black dresses and frilly white aprons, were staring through the window of the hotel conservatory.
3. "Who screamed?" Meyer demanded.

   "Oh—C'est moi... It was I," whispered one of the girls breathlessly. "Look in there" she pointed through the glass.
4. Meyer moved quickly to the window. There on the floor, clearly visible to anyone who looked in, was a large knife with a viciously curved, badly stained blade and a stout wooden handle. Just slightly beyond and all but hidden behind a cluster of philodendrons and a dwarf banana tree, the feet and lower legs of a person were visible. Both heels were in contact with the floor, but the right foot had flopped over on its side. A straw hat could be seen on the floor just inside the door and beside it an open lunch basket. Meyer looked at his watch.

   "12.20," he said almost to himself.

5. He rattled the doorknob but was not surprised to find it locked. A quick check showed it was bolted from the inside. Turning to the ladies he said, "Pardon Mesdames, but is there any other door into the conservatory?"

   "Non, Monsieur, none," replied the more composed of the ladies.

   "Which of you was the one to witness this affair?" asked the great detective.

   "Ah!" replied the older of the two, "we were just passing the doorway on our way to lunch when I glanced in the door and— there you see it Monsieur."

   "Do you know who that can be?"

   "Why, it can only be Jules, the gardener. The work pants; the heavy boots with dirt and grass stuck to the soles. See, even his hat here by the door. But, why, why would anyone... . . . ," she left the rest of the sentence unspoken, her eyes resting on the brutal weapon on the floor.

6. "Hmmmmm!" mused the intrepid investigator, "Clearly an interesting problem for me to solve. The door locked from the inside. No other exits. No one visible through the glass except the gardener."

   The two girls fluttered beside him, clearly agitated by the whole affair. Meyer himself rubbed his brow gently as if to erase the headache from the previous night's Film Festival showing. He allowed his eyes to wander over the grounds of the Grand Hotel. Beautiful flowers abounded everywhere. Clearly the gardener knew his work well and was able to grow things with ease.

7. His gaze returned to the ladies. "You saw no one leave the greenhouse?"

   "Non, Monsieur."

   "He was not like this when you went to lunch?"

   "Non, Monsieur. He was not like this 15 min ago when we passed this way."

   "When did you last see him?"

   "He was here when I arrived this morning at 7 o'clock. He was almost finished trimming the large hedge near the front of the hotel."

8. "Aha!" exclaimed the super sleuth. "As I suspected from the beginning. I shall have a solution to the problem in a moment." With that he reached into his pocket and brought out . . .

Questions

1. What do you think Earl N. Meyer brought from his coat pocket?
2. Where do you think the incident took place? What is your evidence?
3. What conclusion had the two young ladies come to by the end of

Part 5? On what evidence did they base their conclusions?

4. What conclusion had Earl N. Meyer come to by the end of Part 5? What evidence do you have to support your answer to the above question?
5. What is your explanation for what has been observed in the drama described here?

Here is our ending to the story. See if it fits the facts as well as your ending did.

. . . his police whistle. Placing the whistle to his mouth he gave a mighty blast into the noontime air.

"Monsieur, please," pleaded the prettier of the two girls. "Have you no respect? You will wake the dead!"

"Precisely my plan," chortled Meyer. "Voilà!" With that he pointed inside the conservatory. Both girls turned to see the body of Jules sitting up and rubbing the sleep from his eyes after his noontime nap. Meyer motioned him to come to the door.

"What can I do for you Mr. Meyer?" he asked through the now open door.

"The ladies were concerned for your health when they saw you on the floor."

"Thank you for your concern," said the gardener. "You see, I have been working very hard to get the grounds ready for the Film Festival. In fact, I arose especially early this morning to trim the hedge before the general public was up and about. I'm afraid my lack of sleep caught up with me after my lunch as I sat in the warm greenhouse. A thousand pardons if I have frightened you."

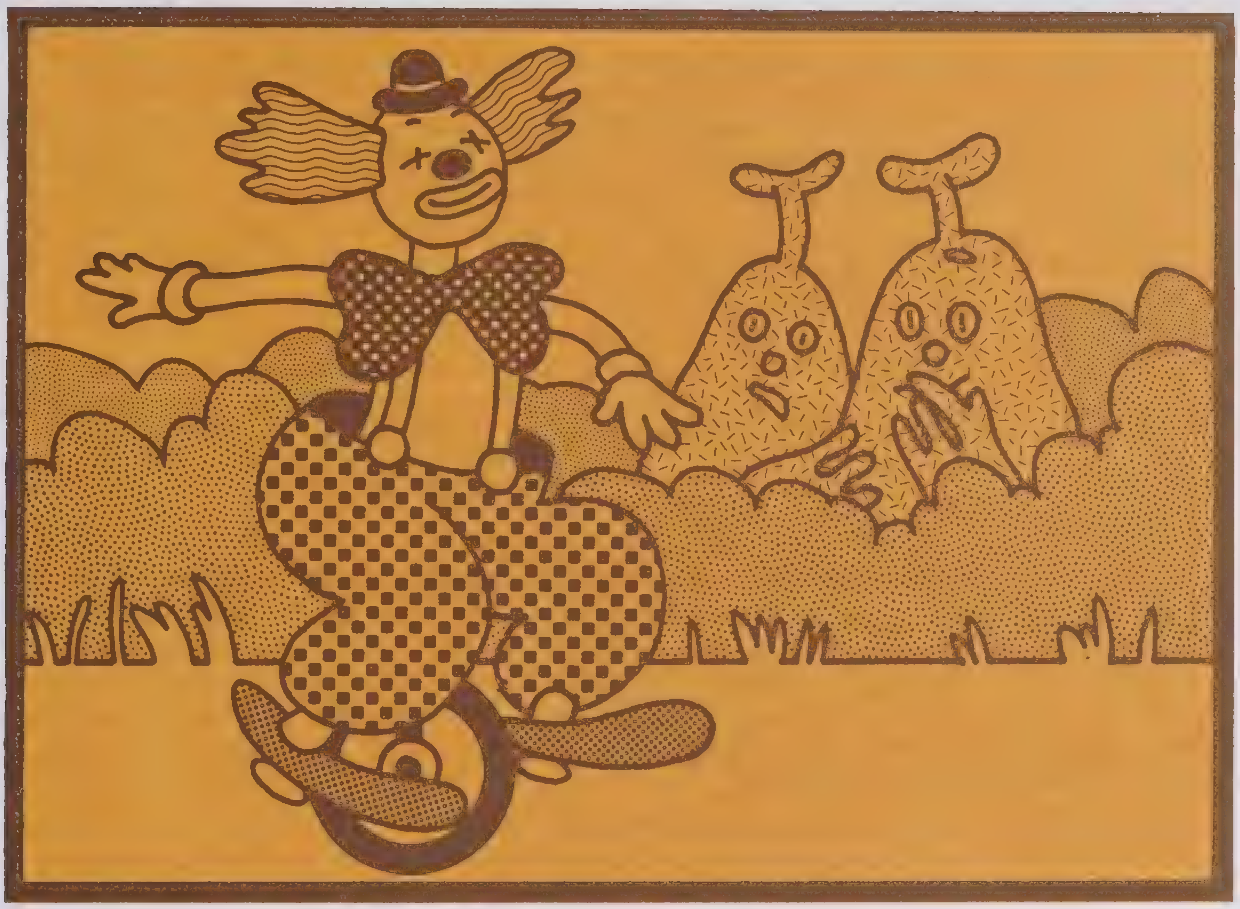
"It was as I suspected," said Meyer. "As in many cases both in crime and in science a conclusion was drawn before all the evidence was gathered. We must all remember that a conclusion based on only part of the evidence is worse than no conclusion at all."

Hypothesizing

Visitors from Outer Space

If you are going to succeed in this course you must allow your imagination to run free at times. Many concepts in chemistry are mental pictures that can only be seen with your imagination. Consider, for example, the following science fiction story.

1. On May 24 of last year, earth was visited for the very first time by beings from another world. Being cautious by nature, these beings chose first to make some observations of the human world before making actual contact. They chose as their obser-



vation site a grove of trees overlooking a major highway that bordered a lake, very near an airport and a railroad line. Their nighttime arrival was carefully selected so that they would remain unnoticed and while they were occasionally rewarded by lights moving through the darkness they actually saw very little until the sun began to light the new day.

2. The first moving object that they saw was a four-wheeled, red, low-slung vehicle which carried two humans in it. This was followed shortly by a larger blue vehicle with many windows, four wheels and only one human, although it could obviously have carried seven or eight. Moments later a green vehicle, again with four wheels, this one with no top that carried four humans — two in front and two behind — went by. Some time passed before a black and white four-wheeled vehicle roared past — a red light flashing on the roof and a terrible wailing sound emanating from it. It contained three humans.
3. At this point the visitors decided to compare notes and find out what they had learned. They concluded that while earth vehicles could come in any colour, be of any shape and obviously different sizes; while these vehicles could have many windows or just a few, carry many people or just one, they all have one thing in common:

Earth vehicles have four wheels.

4. Hearing the sound of an approaching vehicle they turned again to their task — past them sped an orange vehicle with four humans, windows and most important *four* wheels. Feeling very pleased with themselves they awaited the next vehicle. When it appeared they saw an enormous thing — really two things fastened together. The smallest part was in front — it had windows and two humans in it — the larger part was behind and shaped rather like a box, but the wheels! Eighteen all together! To make matters worse the next vehicle had no windows, no roof, one human and two wheels!
5. Feeling very discouraged, they decided to review their observations — suddenly they brightened! Of course, they had been wrong in their initial conclusion because they had said, "Earth vehicles have four wheels," when they should have concluded more simply:
   Earth vehicles have wheels.
6. Suddenly their attention was drawn to the two strips of metal laid side by side and stretching far off into the distance. Coming into view was an enormous thing — roaring and smoking, humans in front, multicoloured and seemingly made of an endless row of boxes — each with eight wheels and each joined to the next in a great long chain! Hundreds of wheels, just the proof they needed.
   Earth vehicles have wheels.
7. Imagine their consternation when in the next few moments they saw out on the water a large thing with hundreds of people, much smoke, many windows, flags, banners and not a single wheel; up in the air a great silver thing trailing smoke, with windows, a long narrow body and two strips of metal stuck out from the sides and not a wheel in sight. Finally, an almost crushing blow, from behind them they heard a roaring clanking sound and turned to see a large yellow thing pushing a mound of earth toward them. The human sat on top, no windows at all, smoke pouring from a pipe sticking up on the top and moving on two large steel tracks — one on each side — that allowed the thing to crawl along the ground.
8. Fearing for their lives the visitors jumped into their spaceship and headed for home. On the way home they had plenty of time to consider what they had seen and to reconsider their generalizations.

Problems

1. Was the conclusion drawn in Part 3 a valid one?
2. Would you have drawn a different conclusion by the end of Part 3?
3. What effect does the information gathered in Part 4 have on the initial conclusion?
4. Is the conclusion drawn in Part 5 a valid conclusion?

5. Does the second conclusion fit all the observations made up to Part 5?
6. What revision do you suppose the visitors made to their conclusion on the way home?
7. If the visitors had stayed longer, what new information might they have gathered about earth transportation that could have helped them?

Physical Change

When you first observed a burning candle, among the most apparent observations were:

1. the burning candle gives off heat;
2. a liquid formed in the bowl of the candle;
3. the liquid in the bowl solidified as it cooled.

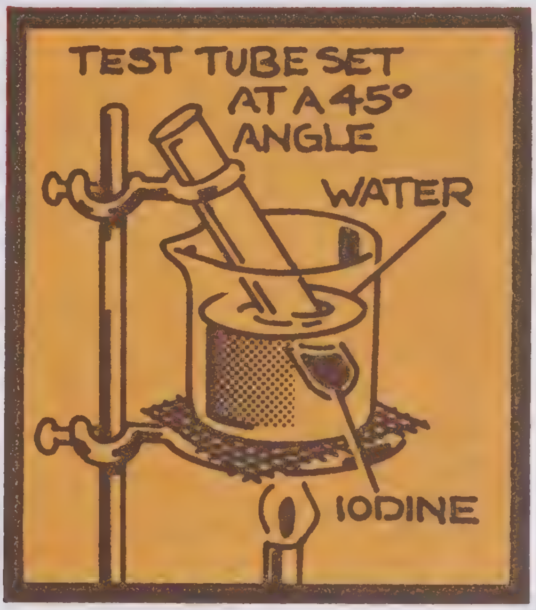
You have undoubtedly noticed that this change from solid to liquid and vice versa is associated with energy or heat. Consider the purpose of a welding torch, the effect of the sun's energy on spring ice, the effect on your hand of grease spattering out of a frying pan. The question of energy and its effect on matter falls into two categories:

1. What is the relationship between matter and energy?
2. How much energy is involved?

The next few experiments will help to answer these questions.

1.5a Heating Iodine

retort stand and ring
clamp
wire gauze
bunsen burner
test tube
beaker
iodine



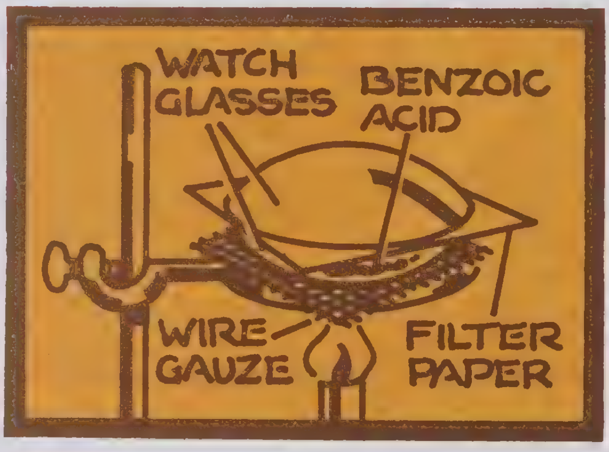
1. Place a couple of very small crystals of iodine in the bottom of a clean dry test tube. Set up the test tube as shown in the diagram.
2. Heat the water with your bunsen burner. Watch the iodine crystals in the test tube. Carefully observe the entire length of the test tube.

Questions

1. Did the original iodine crystals melt in the test tube?
2. Describe what has happened to these crystals.
3. Explain the observations that you made of the iodine crystals.
4. The change that the iodine has undergone is called *sublimation*. Define sublimation.
5. Is sublimation a physical change? Why?

1.5b Heating Benzoic Acid

retort stand and ring
wire gauze
bunsen burner
filter paper
2 watch glasses
spatula
benzoic acid



1. Put as much benzoic acid on a watch glass as will cover the end of your spatula.
2. Cover the watch glass with a filter paper then a second watch glass as shown in the diagram.
3. Support the watch glass "sandwich" on a retort ring and wire gauze.
4. Heat gently with a very low flame for about five minutes.
5. Allow the watch glasses to cool; then carefully remove the upper watch glass and filter paper.

Questions

1. What has happened to the acid crystals that were in the lower watch glass?

2. How do you account for what has happened to these crystals?
3. Name the change that the acid crystals underwent.
4. Was this another example of a physical change? Explain?

States of Matter

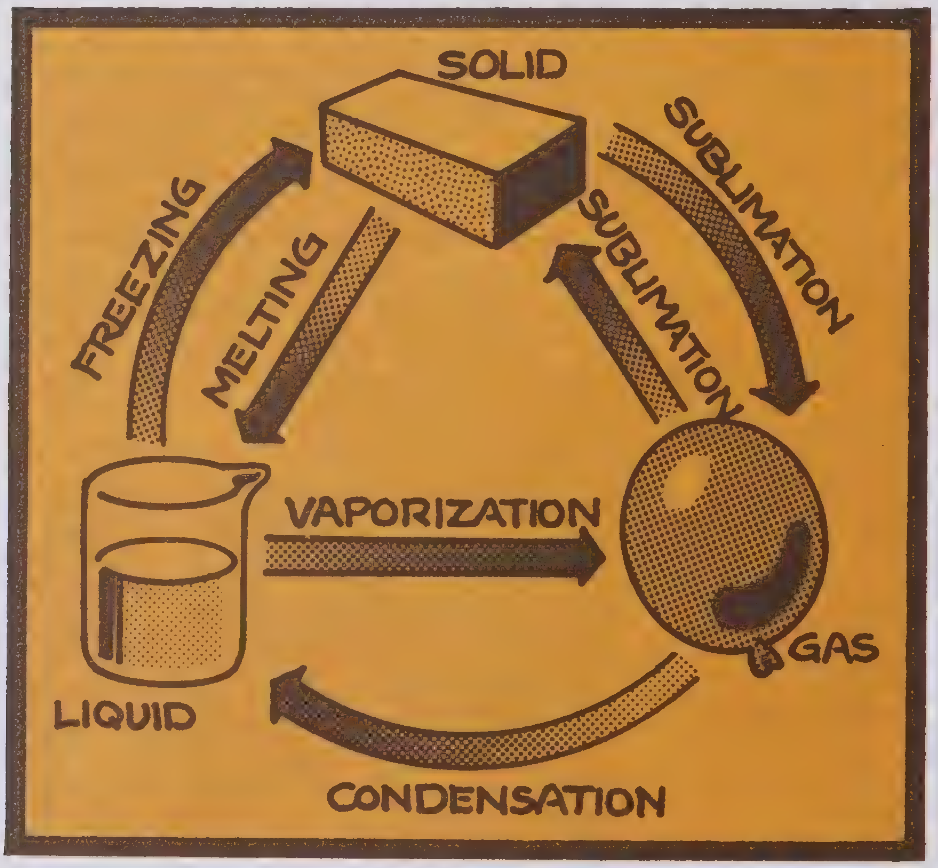
In our last few experiments we have seen matter in several different states. For example, we saw candle material in solid form, as a liquid and perhaps even as a vapour; we saw iodine as a solid, as a vapour and perhaps some of you saw it as a liquid.

We also found that changes from one of these states to another is brought about simply by changing the condition under which the substance is maintained. By a reversal of conditions, matter returns readily to its original form. It does not lose its identify by a change of state. Changes like these are called *physical changes*.

We can identify physical changes by looking at the properties of the substances involved. If a substance goes through a series of changes but, after the final change is completed, still has the same colour, odour, density, melting point, freezing point, solubility in water and so on it is probably the same substance. These changes would be called physical changes. A good example of a physical change is when a tree is turned into sawdust. Although it looks

quite different, sawdust is still wood. We identify a physical change by looking at the properties involved; if the properties indicate that the substance remains the same then the change is said to be a physical change. All changes of state are physical changes. Ice, water and steam are the same substance in different forms.

Before proceeding, however, let us consider other changes of state not yet encountered in our experiments. This perhaps can best be represented by the *change of state triangle*. What change(s) in conditions must exist to bring about each of the changes of state shown in the diagram?



Chemical Changes

There are surely a number of questions that have come to your mind from the discussion of changes of state and physical change. You are probably wondering if all physical changes are changes of state or if all changes in matter are in fact physical changes.

As always we can best get an answer to these two questions by performing more experiments involving change in matter. Try this experiment.

1.6 Other Changes

platinum wire
tongs
2 test tubes
magnesium ribbon
potassium iodide solution
lead nitrate solution
copper wire

Part A

1. Examine and describe a clean platinum wire. Hold the wire in a bunsen flame for about 30 s.
2. Remove the wire from the flame and allow it to cool. Examine and describe the wire now.

Questions

1. Has your description of the wire changed as a result of the heating process?
2. Were the changes that the wire underwent physical changes? Explain.
3. Were the changes state changes or do we already have an answer to our first question?

Part B

1. Repeat Step 1 using a short strip of clean magnesium ribbon (about 3 cm).
   Hold the ribbon with your crucible tongs.

Questions

1. Has your description of the magnesium changed as a result of the heating process?
2. Were the changes that the magnesium underwent physical changes? Explain.

Part C

1. Obtain 5 ml of potassium iodide solution in one test tube and 5 ml of lead nitrate solution in a second test tube. Describe each; then pour the potassium iodide solution into the lead nitrate solution. Describe what happens.

Questions

1. Has your description of the solutions changed as a result of the mixing? Is the change a physical change? Explain.

Part D

1. Repeat Part A using a piece of clean copper wire.

Questions

1. Obviously all changes in matter cannot be classified as physical. The new type of change discovered in this experiment is called a *chemical change*. Clearly distinguish between a physical and a chemical change. In your distinction discuss the permanence of a chemical change compared to a physical change.

Calorimetry

You have, no doubt, already observed that energy is involved in many aspects of science. Things are heated in a burner flame. They absorb heat. They get hot. When heat is added to the substance, it is *endothermic*. Sometimes, as was the case with a burning candle, heat is given off. It is *exothermic*.

As often seems to be the case, our knowledge seems incomplete until we can answer the question "how much?" The claims we see on television for a "New Improved Economical Car" take on a different meaning when we hear that it will run at 1.2 km/ℓ of gasoline and the initial cost of the car will be $12 000.

Chemists are frequently interested in measuring how much heat is involved in a particular reaction. These measurements are usually done with a piece of equipment called a *calorimeter*. The name comes from the word calorie, a unit of heat used in the English system of measurement. Perhaps we should rename the equipment a "jouleimeter," since we will be using the joule (J) as our unit of energy.

First, here is a very brief review of heat measurement. You may remember the following ideas from previous science courses.

1. Energy can neither be created nor destroyed, but it can be passed from one substance to another and it can be changed from one form to another.
2. Heat can only flow or pass from warm to cool. In a sense heat can only go "downhill" in terms of temperature.
3. A combination of these two ideas leads to the idea that the heat lost by a warm body is equal to the heat gained by a cool body. Herein lies the key to calorimetry.

If the heat produced by a chemical reaction is used to warm some other substance, for example, water, then the heat gained by the water will be equal to the heat produced by the reaction. Some time ago you learned how to determine the heat gained or lost by water, using the relation:

$$\text{Heat} = \text{mass} \times \text{change in temp.} \times \text{specific heat capacity}$$

Mass is measured in kilograms; temperature is measured in degrees Celsius; specific heat is measured in joules per kilogram kelvin (J/kg · K).

This will give heat measured in joules.

Probably you learned to do this calculation using calories. The only change is in the units of specific heat capacity.

The specific heat capacity of water is 1 cal/(g·°C).

This is the same as 4200 J/(kg·K).

Remember to change the Celsius scale to the kelvin you just add 273. Freezing point on the Celsius scale is 0°C. On the kelvin scale it is 273 K.

Let us now look at a couple of worked examples to refresh your memory.

Example 1

10 g of water at 10°C are warmed to 20°C. How much energy is required?

Before going any further, write out the information given in the question.

Mass of water	10 g	(m)
Starting temperature	10°C= 283 K	(t_1)

Final temperature	20°C = 293 K	(t_2)
Change in temperature	10 K	$(\triangle t)$

We also know the specific heat capacity of water:

$$\begin{aligned}\text{Specific heat capacity} &= c \\ &= 4200\ \text{J/(kg·K)}\end{aligned}$$

And we know the relationship:

$$\text{Heat} = \triangle t \times c \times m$$

Note that the mass is given in grams, whereas we need it in kilograms, but 1 kg = 1000 g.

All that is required now is to substitute the values for each variable into the equation.

$$\text{Heat} = 0.01\ \text{kg} \times 10\ \text{K} \times 4200\ \text{J/(kg·K)}$$
$$\text{Heat} = 420\ \text{J} = 4.2 \times 10^2\ \text{J}$$

Therefore 4.2×10^2 J of energy are required to warm 10 g of water through 10°C. If we are going to cool 10 g of water through 10°C we would have to remove 4.2×10^2 J of energy

Example 2

How much energy would be required to heat a 1 kg block of steel at 0°C to 500°C, given the specific heat capacity of steel is 400 J/(kg·K)?

Once again, write down what we are given:

Mass of steel	1 kg	(m)
Initial temperature	0°C = 273 K	(t^1)
Final temperature	500°C = 773 K	(t_2)
Specific heat capacity	400 J/(kg·K)	(c)

The change in temperature is $t_2 - t_1$

$$773\ \text{K} - 273\ \text{K} = 500\ \text{K}$$

Now we are ready to substitute into the relationship.

$$\begin{aligned}\text{Heat} &= m \times \triangle t \times c \\ \text{Heat} &= 1\ \text{kg} \times 500\ \text{K} \times 400\ \text{J/(kg·K)} \\ &= 200\,000\ \text{J} = 2 \times 10^5\ \text{J}\end{aligned}$$

Therefore 2×10^5 J of energy are required to heat 1 kg of steel from 0°C to 500°C

Now for the calorimeter. The heat produced by the reaction is used to warm some water. The water is held in some sort of container. This means the container gets warm as well, so we must account for *two* quantities of heat:

1. the heat absorbed by the water;
2. the heat absorbed by the calorimeter.

Assumption

Unless you are told differently, assume that the container and the contents are at the same temperature.

Example 3

500 g of water at 10°C are warmed in a 100 g aluminum can to a temperature of 50°C. Calculate the energy input.

Given
m water = 500 g
m aluminum = 100 g
t_1 water = 10°C = 283 K
t_2 water = 50°C = 323 K
$\triangle t$ = 40 K

From Table 1.2 on page 49 we can get
c water = 4200 J/(kg·K)
c aluminum = 880 J/(kg·K)

Now remember there are two quantities of heat involved — the heat for the water and the heat for the can.

Let's do the water first.

$$\begin{aligned}\text{Heat} &= m \times \triangle t \times c \\ &= 0.5\ \text{kg} \times 40\ \text{K} \times 4200\ \text{J/(kg·K)} \\ &= 84\,000\ \text{J} = 84 \times 10^3\ \text{J}\end{aligned}$$

Therefore 84×10^3 J are required to heat the water.

Now consider the aluminum can. Remember our assumption that the temperature of the can is the same as the temperature of the water.

$$\begin{aligned}\text{Heat} &= m \times \triangle t \times c \\ &= 0.1\ \text{kg} \times 40\ \text{K} \times 880\ \text{J/(kg·K)} \\ &= 3520\ \text{J} = 3.5 \times 10^3\ \text{J}\end{aligned}$$

Therefore 3.5×10^3 J are required to heat the aluminum can. The Total energy is therefore:

$$\begin{aligned}&84 \times 10^3\ \text{J} + 3.5 \times 10^3\ \text{J} \\ &= 87.5 \times 10^3\ \text{J} \\ &= 88 \times 10^3\ \text{J}\end{aligned}$$

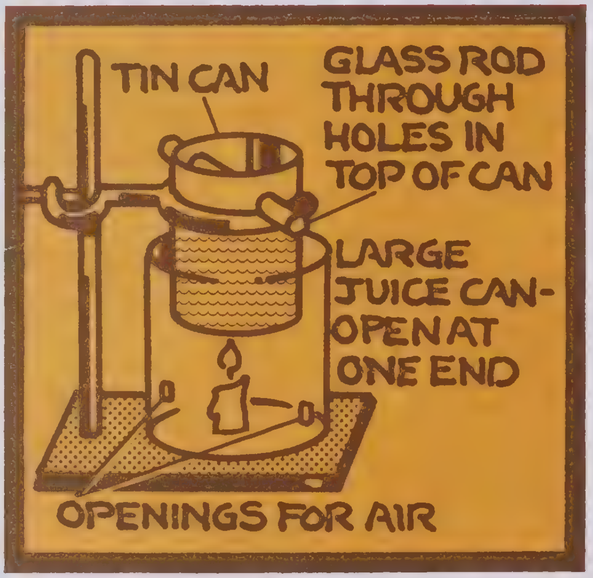
A burning candle gives off heat. Where does the heat come from? We now know that the liquid candle material gives off heat as it solidifies. Is this the full amount or does the burning of the candle material itself produce heat? If there is more than one source, how do the relative amounts of energy produced compare?

In this experiment we shall look for an answer to these questions. We shall determine how much heat is released and how much candle material had to burn or had to solidify to produce that amount of heat.

We must have some means of measuring the heat released and also of determining how much candle material is burning or how much is solidifying. The heat in each case will be given up to some water. We must now decide what we will have to know about the water if we are to determine how much heat is released to it.

1.7a Heat Released by Burning Candle Material

retort stand and ring
tin lid
small tin can
large tin can
glass rod
candle
thermometer
balance


TIN CAN
GLASS ROD THROUGH HOLES IN TOP OF CAN
LARGE JUICE CAN-OPEN AT ONE END
OPENINGS FOR AIR

1. Mount a candle on a tin lid by allowing a few drops of the liquid candle material to fall on the lid where the candle will be placed.
2. Determine the mass of the unlit candle and the lid.
3. Determine the mass of the small metal can.
4. Fill the small metal can about two-thirds full of water that has been previously cooled to about 10°C below room temperature.
5. Set up the apparatus as shown in the diagram.
6. Record the temperature of the cold water.
7. Light the candle and heat the water until it is almost as far above room temperature as it was intially below room temperature.
8. Extinguish the candle flame and then record the highest temperature reached by the water in the can. (Watch your ther-

mometer and stir gently. The water temperature may rise a degree or so after the candle flame has been extinguished.)

9. Determine the mass of the can and water.
10. Determine the mass of the candle and lid.
11. Determine:
    (a) mass change of the candle;
    (b) mass of the water heated;
    (c) temperature change of the water;
    (d) quantity of heat given to the water by the burning candle:
    $$m \times \triangle t \times c \text{ (specific heat capacity)}$$
    (e) heat released to the water per gram of candle burning:
    $$\text{quantity of heat/mass of candle used} = \frac{\text{J}}{\text{g}}$$

Questions

1. Has the mass of the candle increased or decreased?
2. If the mass of the candle has decreased, what has happened to the candle material lost?
3. Discuss the possible sources of error.
4. Decide if the amount of heat released per gram of candle material burning is more than or less than the accepted value.

1.7b

Heat Released by Solidifying Candle Material

test tube with candle wax
test tube
balance
2 250 ml beakers
thermometer

1. Obtain a test tube that is half filled with solid candle material and determine its mass. Determine the mass of an identical empty test tube.
2. Place the test tube in a beaker of hot water until the candle material has completely melted.
3. Determine the mass of a second beaker (about 250 ml) to the nearest 0.01 g. Add cold water to the beaker until it is about two-thirds full. The water temperature should be about 3°C below room temperature.
4. Remove the test tube and candle material from the hot water and hold it in the air until solidification just begins (cloudiness near the bottom of the test tube will be noticed just as solidification starts).
5. Record the temperature of the cold water and immediately place

the test tube and candle material into the beaker of cold water.

6. Stir the water with the test tube until *all* the candle material has solidified. Record the highest temperature reached by the water in the beaker.
7. Determine the mass of the beaker and water.
8. Determine:
   (a) mass of candle material in the test tube;
   (b) mass of water heated;
   (c) temperature change of the water;
   (d) quantity of heat released to the water in the beaker by the solidifying candle material;
   (e) quantity of heat released to the water for each gram of substance solidifying.

Questions

1. Name the changes of state that the candle material went through in Part B of this experiment.
2. Did the candle material retain its identity throughout these changes?
3. Compare the heat released in Part A with the heat released in Part B.

Problems

1. How much heat would be released if 100 g of glycerol were burned?
2. Assuming that a candy bar is mostly sugar, how much energy is there in an average-sized bar?
3. How much heat would be required to change 10 g of ice at at 0°C into water at 0°C?
4. How much salt could be melted by a 1 g of burning sulphur?
5. A candle made from p-dichloro benzene is used to heat a 500 g sample of water from 20°C to 50°C. What mass of p-dichloro benzene will be consumed in the reaction?
6. In order to determine the heat of combustion of butane, a disposable lighter is used to warm the water in a calorimeter. The following data was collected in the experiment:

   Initial mass of lighter = 54.04 g
   Final mass of lighter = 52.26 g
   Initial temperature of calorimeter = 12.0°C
   Final temperature of calorimeter = 31.8°C
   Mass of aluminum calorimeter = 25.2 g
   Mass of calorimeter and water = 1025.2 g

   Calculate the value of the heat of combustion of butane.
7. An Eskimo mother is melting ice to make seal soup for her

family. She uses a whale oil lamp to supply the heat. How much ice at −20°C could she heat to 90°C if she has 10 g of whale oil available to her? Assume that half the heat produced is lost to the air in the igloo and to the container.

$\triangle H_c$ whale oil = 30 000 000 J/kg

$\triangle H_f$ = heat of fusion measured at 15°C
$\triangle H_c$ = heat of combustion measured at 15°C
$\triangle H_v$ = heat of vaporization measured at boiling point
c = specific heat capacity measured at 15°C
All measured in joules per kilogram (J/kg)

Table 1.2

Substance	$\triangle H_f$	$\triangle H_c$	$\triangle H_v$	c
aluminum	321 000	15 940 000	—	880
carbon dioxide (solid at −50°C)	189 000	—	229 000	750
P-dichlorobenzene	124 000	19 000 000	—	920
glycerine (glycerol)	199 000	18 000 000	—	1370
lead	21 400	904 000	—	130
paraffin wax	147 000	40 500 000	—	2880
sodium chloride (salt)	518 000	—	—	840
sucrose (sugar)	—	16 500 000	—	—
sulphur	55 000	9 200 000	—	750
water	—	—	2 257 000	4200
ice	334 000	—	—	2090

Solutions

Cola, tea, drinking water, washing bleach, household cleaners and gasoline all have something in common. They are all solutions. In previous years at school you have no doubt studied to some extent various solutions that you prepared yourself or, like those listed above, solutions that we encounter in our everyday lives. Let's look more closely at one of these solutions and recall some of the things that we should already know about solutions. Our example will be coffee.

Many people begin their day by plugging in the kettle and putting a spoonful of instant coffee in a mug. When the water is boiling it goes into the mug and with a brief stir away go the coffee grains. They disappear into the water, mixing uniformly throughout the water in the mug. They *dissolve*. The coffee grains in this solution are referred to as the *solute* (that which is dissolved) while the water is called the *solvent* (that which does the dissolving). Often

Alkalai lake, Saskatchewan. The evaporation of water from shallow lakes leaves behind a deposit of the dissolved minerals. This photograph shows a large deposit of minerals formed in one of these "alkali lakes." *Courtesy of Department of Tourism and Renewable Resources, Saskatchewan.*

solutions consist of more than one solute. Our cup of coffee is an example. Probably most of us add a little cream as well as several cubes of sugar, both of which disappear into the solvent. Some people like a lot of water and a small amount of coffee. This would be a dilute solution. Others prefer a concentrated solution, which would mean less water and more coffee.

Let's go back for a moment to the original coffee solution and look more closely at it and some of its properties as a typical solution.

First of all, the solid particles disappear, distributing themselves evenly throughout the solvent. The entire solution appears the same throughout or is homogenous with *one* phase visible. The solution is coloured, but if a sample is poured into a test tube we will find that, although colour has been imparted to the solution, it is still clear with only one visible phase. This is a property of liquid-solid solutions. The large coffee grains have disappeared from the bottom by breaking up into tiny particles that are so small they may even pass through a filter. At this point look more closely at other liquid-solid combinations to see if they behave in a similar manner to the coffee in hot water.

1.8 Mixing Liquids and Solids

test tubes
test tube rack

sodium chloride
methanol
potassium permanganate
iodine
mineral oil
sodium hydroxide
ammonium nitrate

Table 1.3

Substances	*Evidence of Dissolving*	*Number of Substances Visible*	*Clarity Colour*	*Temp. Change*	*Solution or Mixture*
A small scoop of sodium chloride and 1/2 test tube of water					
A small scoop of sodium chloride and 1/2 test tube of methanol					
1 crystal potassium permanganate and 1/2 test tube of water					
1 crystal iodine and 1/2 test tube of water					
1 crystal of iodine and 1/2 test tube of methanol					
1/4 test tube of water and 1/4 test tube of methanol					
1/4 test tube of water and 1/4 test tube of mineral oil					
5 pellets of sodium hydroxide in 1/2 test tube of water					
A small scoop of ammonium nitrate in 1/2 test tube of water					

1. In this experiment you are to bring the substances indicated in Table 1.3 together in a test tube. In general use about half a test tube of the liquid and a small quantity of the other material.
2. When the substances are in contact shake them well for a moment, then place them in a test tube rack.
3. After settling, examine the mixture and record the necessary observations according to Table 1.3.
4. When you are finished, flush the contents of each test tube down the sink or dispose of them as your teacher directs.
5. In the final column record the conclusions that you reached about the nature of the contents of the test tubes.

Questions

1. How can you tell which combination of substances produced *mixtures*?
2. How can you tell which combination of substances produced *solutions*?
3. Which test tubes, if any, contained both a solution and a mixture?
4. What name is given to the substance that dissolves or seems to disappear in the solution?
5. What name is given to the substance that does the dissolving or seems to remain in a solution?
6. What do we mean when we say something has "dissolved" in something else?
7. Compare a solution and a mixture under the following headings:
   (a) number of components;
   (b) number of phases;
   (c) energy exchange or formation;
   (d) proportion of components.

Mechanical Mixtures

Homogeneity, the condition of being homogeneous or uniform throughout, is a property that could describe at least six of the combinations of materials that you prepared in Experiment 1.8. The other three combinations are better described by the term *heterogeneity*. They are equally important in classifying matter. The combinations that were heterogeneous have parts that are distinctly visible or become visible after some operation is carried out on them. For example, a rock called porphyry consists of quartz, feldspar, pyroxene, hornblend and mica. When observed closely, large crystals of one mineral can be observed embedded in a finer-grain mass. Tomato juice placed into a centrifuge for a few minutes separates out into a pinkish liquid and a deep red pulpy layer. Substances like the two described above are termed *mechanical mixtures*. Both solutions and mechanical mixtures are of the utmost importance and concern to us because many of the things that we use or eat everyday are solutions. In the preparation of these things we must become familiar with the properties of solutions, the effectiveness of various solvents and the effect that solutes may have on these solvents.

On the other hand, many of the valuable minerals that we use are found as mechanical mixtures like the minerals in porphyry, or are found in solution form like the many materials that are recovered from petroleum. For the gasoline in petroleum or the mica in porphyry to be of any value to us they must first be separated from either the mechanical mixture or the solution in which they are found.

1.9

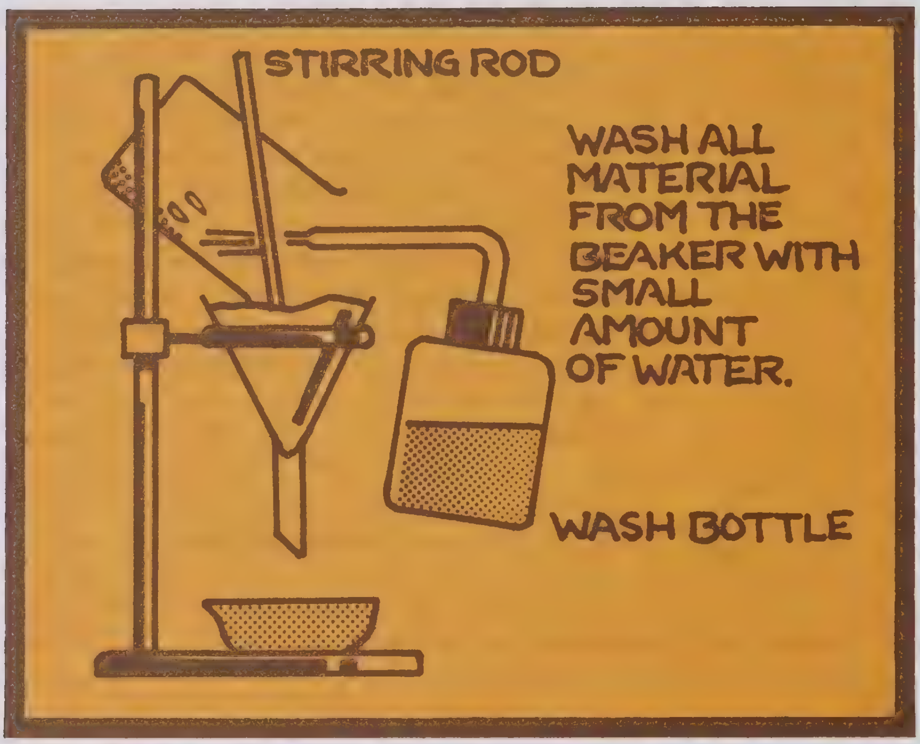
Separating the Constituents of Mechanical Mixtures and Solutions

powdered calcium carbonate
sodium chloride (fine)
filter paper
3 watch glasses
water
spatula
wash bottle
stirring rod
evaporating dish
150 ml beaker
bunsen burner

1. Put about 5 g of powdered calcium carbonate and 3 g of fine sodium chloride on separate dry watch glasses. Examine each of the substances carefully and record *all* observations.

Note: The amounts selected do *not* have to be exactly those suggested. Just be sure that you know the exact amount of each.

2. Mix the sodium chloride into the calcium carbonate. Use your spatula or scoopula to do the mixing and continue to mix until the two substances are combined uniformly. Record your observations.
3. Place the mixture into a clean 150 ml beaker and add 25 ml of distilled water. Stir well.


STIRRING ROD
WASH ALL MATERIAL FROM THE BEAKER WITH SMALL AMOUNT OF WATER.
WASH BOTTLE

4. Determine the mass of a filter paper and of an evaporating dish. Filter the mixture that you now have in the beaker and collect the filtrate in the evaporating dish. Use *small* amounts of distilled water to wash *all* materials from the beaker into the filter.
5. When all the liquid has passed through the filter, heat the evaporating dish until all the water has evaporated. As spattering may occur especially near the end, the heating should be done carefully.
6. Allow the evaporating dish to cool. Describe the contents. Determine the mass of the dish and contents. Taste a small amount of the material in the dish.
7. Unfold the filter paper and place it on a clean watch glass. Leave it to dry or place it in a drying oven if one is available. In the next laboratory period determine the mass of the dried filter paper then discard it.

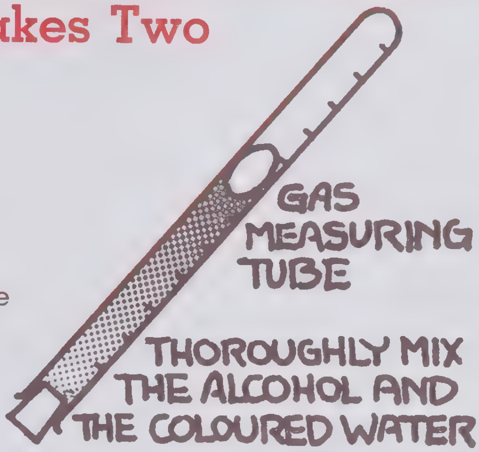
Questions

1. What property(s) of the constituents of the mixture enabled you to separate them by first adding water and then filtering?
2. What property of water were you using in recovering the sodium chloride from the filtrate?
3. Knowing the original and final masses of the two constituents of the mixture, calculate the percentage by mass recovered for each constituent. Account for any similarities or differences in these two percentages.

4. Would the method outlined in this experiment be a reasonable way to separate powdered iron from sand? Explain.
5. In this experiment we prepared a mixture-solution combination. Part of the separation was done by a mechanical process — filtration. List other mechanical processes that may be used to separate the constituents of mixtures. Give one example of a mixture that each process may be used with.
6. The filtrate was a solution. Was the separation of the solvent and the solute of this solution a mechanical process? Explain.

1.10a One and One Makes Two (Demonstration)

methanol
water
food colouring
burette or gas measuring tube



1. Prepare about 50 ml of coloured water by adding a few drops of food colouring. Add exactly 25 ml of the coloured water to a 50 ml gas measuring tube. (Use a burette if such a tube is not available.)
2. Holding the tube at about a 45° angle, carefully add 25 ml of methanol on *top* of the water. The tube will now contain exactly 50 ml of liquid. Record the volume.
3. Cork the tube and invert it several times (seven or eight at least).
4. Hold the tube in a vertical position with the corked end facing upward. Allow *all* liquid to run off the insides of the tube. Record the volume now.

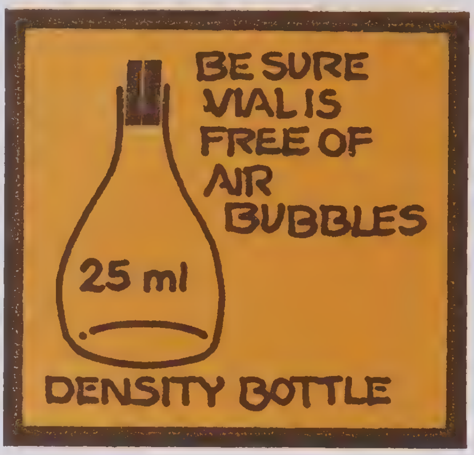
Questions

1. Why was the water coloured?
2. Did the water and alcohol form a solution when they were first put together? Explain.
3. Did the water and alcohol form a solution after the tube had been inverted several times? Explain.
4. Did any of the liquid escape from the tube?
5. Can you offer an explanation for the observations made on the volume in this demonstration? (If you are unable to, don't worry. As we gather more information about the makeup of matter an explanation will become obvious.)

1.10b

Is the property of solutions observed in the previous demonstration unique to liquid-liquid solutions? Try this experiment and find out.

potassium chloride
distilled water
balance
graduated cylinder
thermometer
1 density bottle (25 or 50 ml)
125 ml beaker


BE SURE VIAL IS FREE OF AIR BUBBLES
25 ml
DENSITY BOTTLE

1. The density of potassium chloride is 1.984 g/cm³. Calculate the mass of 5 cm³ of potassium chloride to the nearest 0.01 g. Transfer it all to a 125 ml beaker.
2. Measure *exactly* 50 ml of distilled water at 20°C and add it to the beaker. Dissolve the potassium chloride and again record the temperature. (If it is not 20°C, warm it slightly with your hands or a water bath.)
3. Determine the mass of the density bottle, and when the temperature of the solution is again at 20°C, fill the density bottle with it. Insert the capillary stopper with light pressure. Make sure that there are no air bubbles present. Dry the outside of the bottle and then determine the mass again.

Observations and Calculations

1. Mass of 5 cm^3 of potassium chloride ________________ .
2. Mass of 50 ml of distilled water ($D = 1\ \text{g/cm}^3$) ________________.
3. Total volume = 55 ml Total mass = ________________.
   Theoretical density of potassium chloride solution
   $= \frac{\text{mass}}{\text{volume}} =$ ________________.
4. Mass of the density bottle = ________________ .
5. Mass of the density bottle and solution = ________________ .
6. Mass of the solution = ________________ .
7. Volume of solution (from density table) = ________________.
   Actual density of potassium chloride solution
   $= \frac{\text{mass}}{\text{volume}} =$ ________________.

Questions

1. Compare the theoretical density of the solution to the actual measured density. Are they the same? If not, which is the largest?
2. A difference in the two densities *must* mean a change in either the *mass* or the *volume* as the solution was formed. Can you account for any change in mass during the formation of the solution?
3. What conclusion can we draw about the formation of the potassium chloride *solution*?

Properties of Matter

If you ponder over your day you may well be able to think of a number of solutions or mechanical mixtures that you have encountered or made use of. The list could be almost endless. For example, milk, gasoline, orange juice, soft drinks, soap, rocks or pavement may be either solutions or mechanical mixtures. This is one reason for much of the interest and concern about them and new ones yet to be discovered. In many instances, the things that we use are found either as a solution or as a mechanical mixture. This means we have to develop means of separation. If we think back to the separation of sodium chloride and the carbonate, a difference in *properties* was necessary before the two substances could be separated. In this example, the fact that one of the constituents dissolved in water and the other did not enabled us to separate them. Separations of matter are always based on some difference in the properties of the substances involved.

Let's consider a few of the examples listed above and look at the properties that are used in separating them. Milk, for example, contains about 4 percent butterfat which can be used to make other products if separated from the rest of the milk. These fat globules have a different *density* than the rest of the milk and therefore slowly rise to the top. This process is very slow so a machine that works on the same principle as a centrifuge is used to speed up the separation. Since the fat in the milk is a mechanical mixture, a mechanical means can be used to separate the two constituents.

Gasoline is not found as such in nature, but is part of crude oil or petroleum. The gasoline is dissolved in this mixture forming a solution. It cannot, therefore, be separated by mechanical means. How did we separate the salt that was dissolved in the water? We boiled the solution and since the water boiled at a lower temperature than the salt, it evaporated before the salt. If we had wanted to recover the water as well, all we would have had to do was cool the vapours and condense them. The process here is called *distillation*. A solution is distilled first by evaporation and then by condensation. Gasoline is only one of many constituents of petroleum and is recovered by distillation. Since it boils at a different temperature from other constituents it can be cooled and condensed at that temperature.

A comparison of solutions and mixture is given in Table 1.4.

Table 1.4

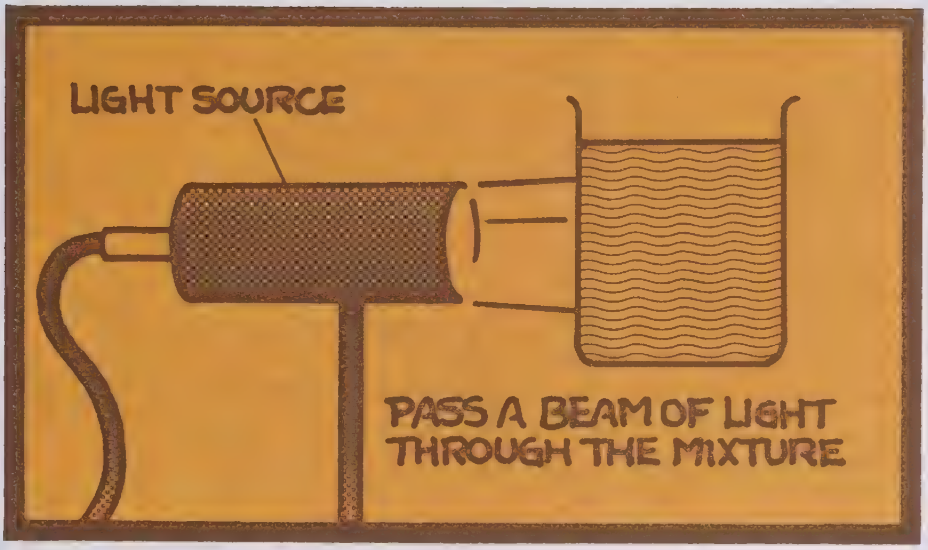
	Mixtures	*Solutions*
Number of Substances	two or more	two or more
Proportions	any proportion	any proportion to solubility
Uniformity	heterogeneous	homogeneous
Number of Phases	two or more	one
Clarity	opaque or translucent	usually transparent
Energy of Formation	none	may be exothermic or endothermic
Properties	no change	some change, i.e. volume reduction

1.11 Colloids

Are all mixtures clearly defined solutions or mechanical mixtures?

150 ml beaker
250 ml beaker
distilled water
0.2 M ferric chloride
bunsen burner
filter paper
dialysis tubing

1. Put 100 ml of distilled water into a clean 150 ml beaker and boil it for several minutes.
2. Add 0.2 M $FeCl_3$ (0.2 Molar ferric chloride) at a rate of 1 drop per second until a deep red suspension is formed.
3. Pass a beam of light through the mixture now in the beaker. A beam of light from a ray box will be suitable. Record your observations.
4. Filter a sample of the mixture. Describe the filtrate. Was there a residue? Again pass a beam of light through the mixture.



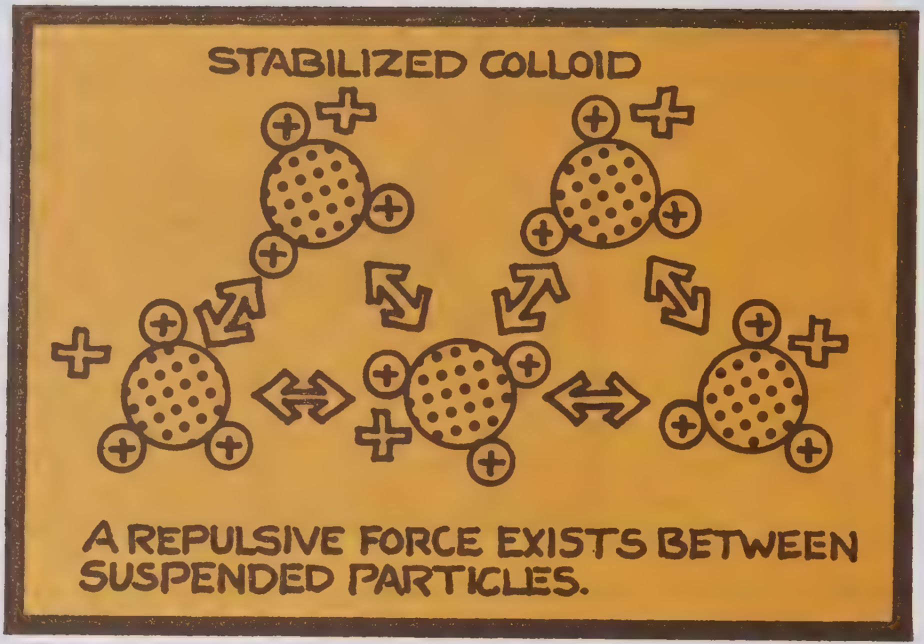
5. Place a sample of the mixture into some dialysis tubing and leave it in a 250 ml beaker of water until the next laboratory period. Record your observations.

Questions

1. Was the mixture in the beaker transparent? Describe the observation(s) on which you base your answer.
2. Did solid particles settle to the bottom of the beaker?
3. Did filtering separate the suspended particles from the mixture? Was the filtrate transparent?
4. Explain the observations made in Part 5 of the experiment.
5. Were the properties of this mixture characteristic of either solutions or mechanical mixtures? Explain.

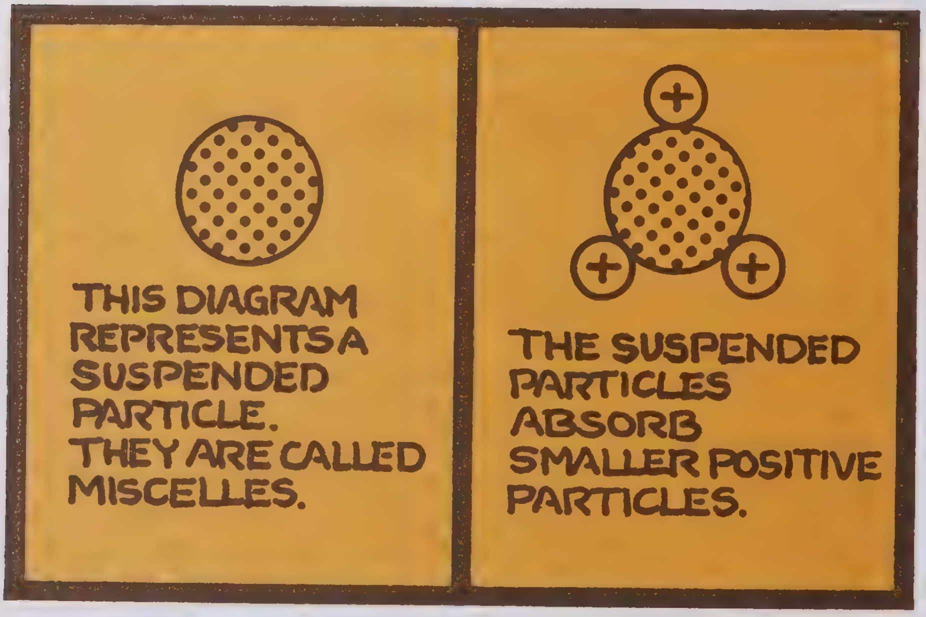
Colloidal State

The term *colloidal state* refers to a dispersion of one type of particle into another. A colloidal dispersion is not a true solution. The particles of a colloid are somewhat larger than those found in a true solution. The particles of a colloid, although much larger than those of a solution, are still small enough to pass directly through a filter. This makes these suspended particles different from those found in a mechanical mixture even though the suspended particles do form a distinct phase. One of the most distinctive properties of colloidal particles is their tremendous surface-to-mass ratio. What this means is that there is a great amount of surface area exposed for the mass of the particle.


STABILIZED COLLOID
A REPULSIVE FORCE EXISTS BETWEEN SUSPENDED PARTICLES.

The particles of a colloid do not settle out like those of a mechanical mixture for several reasons. One of these if that they are continually being jostled and knocked about by other particles which does not allow them to settle. A second reason is that charged particles (of the same charge) are *adsorbed* into the surface of the suspended particles, making each particle a charged clump. Each of these clumps has the same charge, and they therefore repel each other. As this force of repulsion is greater than the pull of gravity, the colloidal particles remain suspended.

Some colloids are extremely stable and last for many years. Michael Faraday prepared colloids of gold in water and now, over a hundred years later, the colloids he prepared are still stable.


THIS DIAGRAM REPRESENTS A SUSPENDED PARTICLE. THEY ARE CALLED MISCELLES.
THE SUSPENDED PARTICLES ABSORB SMALLER POSITIVE PARTICLES.

Types of Solutions

Normally when one thinks of a solution, it is always along the lines of a solid dissolved in a liquid. So far in our study of solutions the emphasis has been on the solid-liquid type of solutions. Solutions may, however, exist in all the possible combinations of a solid, liquid or a gas dissolved in a solid, liquid or gas.

One of the most important properties of metals is that, when melted and mixed together, they will combine evenly to form a homogeneous mixture which, when cooled, will form a solid solution. Solder, for example, is prepared by dissolving molten lead with molten tin in perhaps a 50-50 proportion. Numerous other alloys, such as silver for coins, different kinds of steel, and brass, are solids dissolved in solids, or solid solutions.

Mercury is the only metal that is liquid under normal conditions. It is found to some extent in its pure form in nature but it is usually in the form of a compound called cinnabar, (mercury sulphide). Mercury has a number of unique properties that make it valuable for a variety of uses. For example, it does not wet glass, so it can be used in thermometers or barometers. A very important property of mercury is its ability to dissolve other metals to form solid solutions. These alloys are called *amalgams*. Silver amalgam, for example, is used to fill tooth cavities. Silver, which is sometimes found as an element distributed throughout rock or ore, is concentrated by dissolving it in mercury to form a silver amalgam.

Most other types of solutions need little mention. We are familiar with them if we stop and think for a moment. Gases dissolved in liquids are encountered every time we open a bottle of carbonated drink. Liquid dissolved in gas is, of course, all around us as water dissolved in air. Pure air itself is a solution of gases, essentially oxygen dissolved in nitrogen. Table 1.5 summarizes some of the various types of solutions.

Table 1.5

	Types of solutions
Solid-Solid	Solder (tin and lead)
Solid-Liquid	Coffee (coffee and water)
Solid-Gas	Dry ice in air
Liquid-Solid	Mercury in silver (amalgams)
Liquid-Liquid	Alcohol in water
Liquid-Gas	Humid air
Gas-Liquid	Acetylene in acetone
Gas-Gas	Air tanks (oxygen in nitrogen)

Solubility

Solubility is the amount of a solute (expressed in grams) that will dissolve in 100 g of solvent to make a saturated solution at a certain temperature.

1.12

Solubility of Potassium Chlorate and Copper Sulphate

The last experiment reminded you of some of the things you have learned about solutions in previous years. The next experiment will attempt to broaden your knowledge of solutions.

In this experiment you and your partner will not work together but will share your results. One of you will do the experiment using copper sulphate as the solute; the other will use potassium chlorate.

test tube
stopper
thermometer
copper sulphate
potassium chlorate

1. Measure out exactly 10 ml of cool tap water into a clean test tube.
2. Add a small scoop of your solute, stopper the tube and shake well.
3. Allow the tube to stand a moment and record your observations.
4. Add a second small scoop of solute, stopper, shake and allow to stand again.
5. Repeat this procedure until after shaking, some solid solute remains in the test tube.
6. Warm the solution in a water bath to about 40°C. Stopper, shake and allow to stand.
7. Add solute a scoopful at a time until again some remains undissolved.
8. Warm the solution at 60°C. Stopper, shake and allow to stand. Again add solute a scoopful at a time, shaking after each addition until some solid remains.
9. Stopper the test tube and cool it under the cold water tap.

Questions

1. What two observations lead you to conclude that copper sulphate and potassium chlorate dissolve in water?
2. Describe the solution formed in Step 3 in terms of its ability to dissolve more solute. What name is given to this type of solution?
3. What type of solution has been formed by the end of Step 5? Describe this solution in terms of its ability to dissolve more solute.
4. What type of solution is formed as the solution is cooled down? Describe this solution in terms of its ability to dissolve more solute.
5. How could one determine if a given solution was saturated, unsaturated, or supersaturated?

Solubility Curves

Now that you have some qualitative information about dissolving it is time to get some quantitative information about dissolving and solutions. The last experiment showed that a given amount of solvent can only dissolve a certain limited amount of solute. When this maximum point is reached the solution is said to be *saturated*. The next experiment deals with saturated solutions.

For this experiment each group will follow a slightly different procedure. At the end of the period you will share your results with other groups. Each group will prepare a saturated solution of copper sulphate at a different temperature. Your teacher will assign you the temperature for your particular solution.

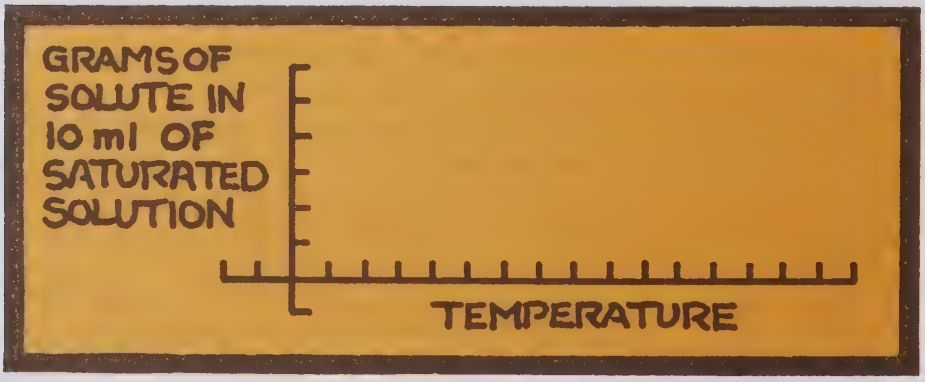
1.13 Preparing Solubility Curves

balance
100 ml beaker
thermometer
25 ml graduated cylinder
evaporating dish
copper sulphate

1. Measure approximately 20 ml of distilled water into a clean dry 100 ml beaker.
2. Determine the mass of a clean dry evaporating dish.
3. Cool or heat the beaker until the water reaches the assigned temperature.

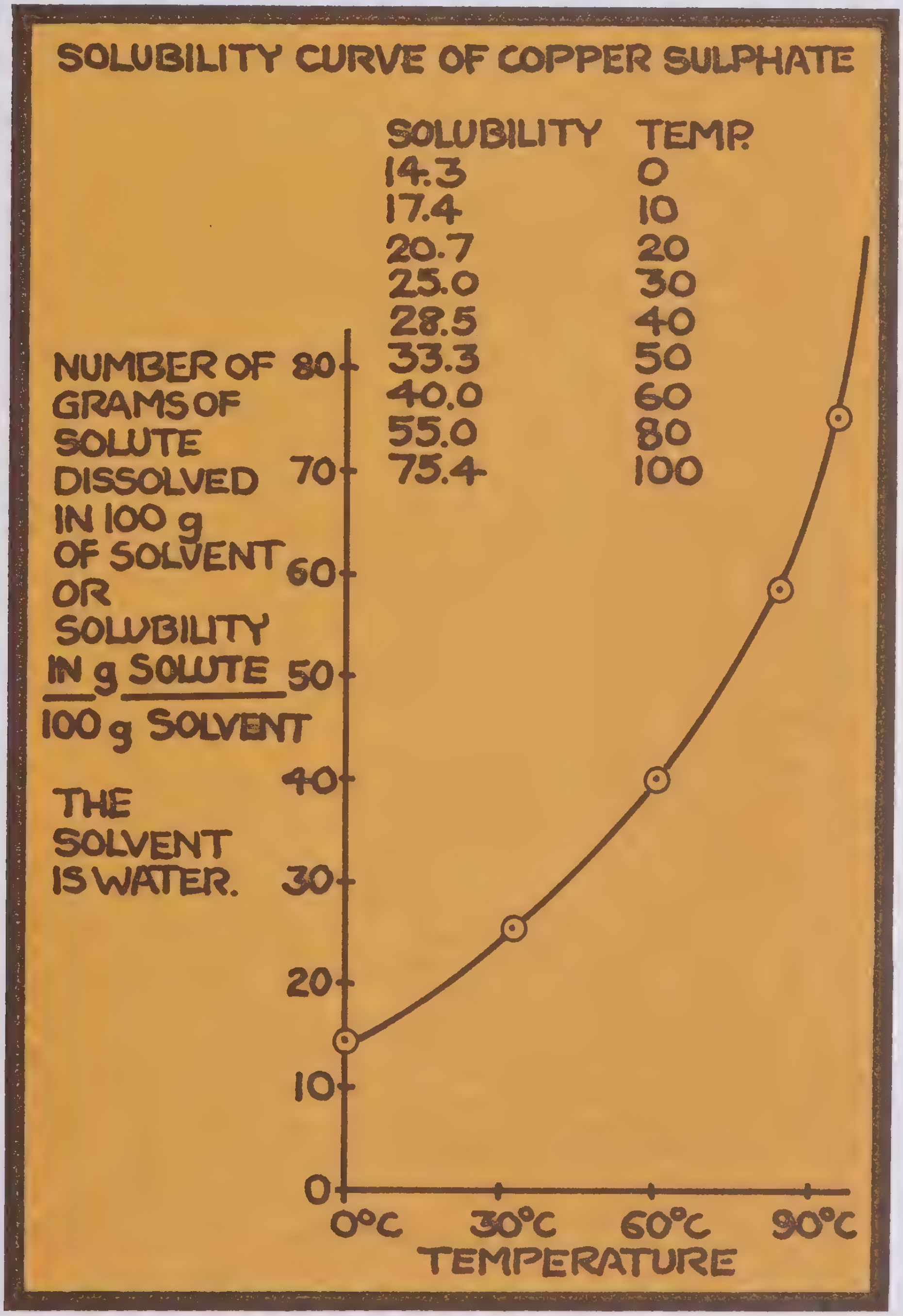
   You must maintain this temperature while you do the next part!

4. Add copper sulphate while stirring until you are sure the solution is saturated, then stir one minute more.
5. Still maintaining the temperature at the assigned value, allow the solution to stand long enough for the suspended solid to settle.
6. Carefully pour exactly 10 ml of the saturated solution into the evaporating dish.
7. Slowly evaporate the solution to dryness. (Something of interest may happen near the end of the heating.)
8. When the residue is dry, allow the dish to cool and determine its mass again.
9. Calculate the number of grams of solute dissolved in 10 ml of saturated solution. Enter this figure in the table your teacher has prepared on the blackboard.
10. Copy down the table. Plot a graph of grams of solute dissolved in the saturated solution against the temperature of the solution.



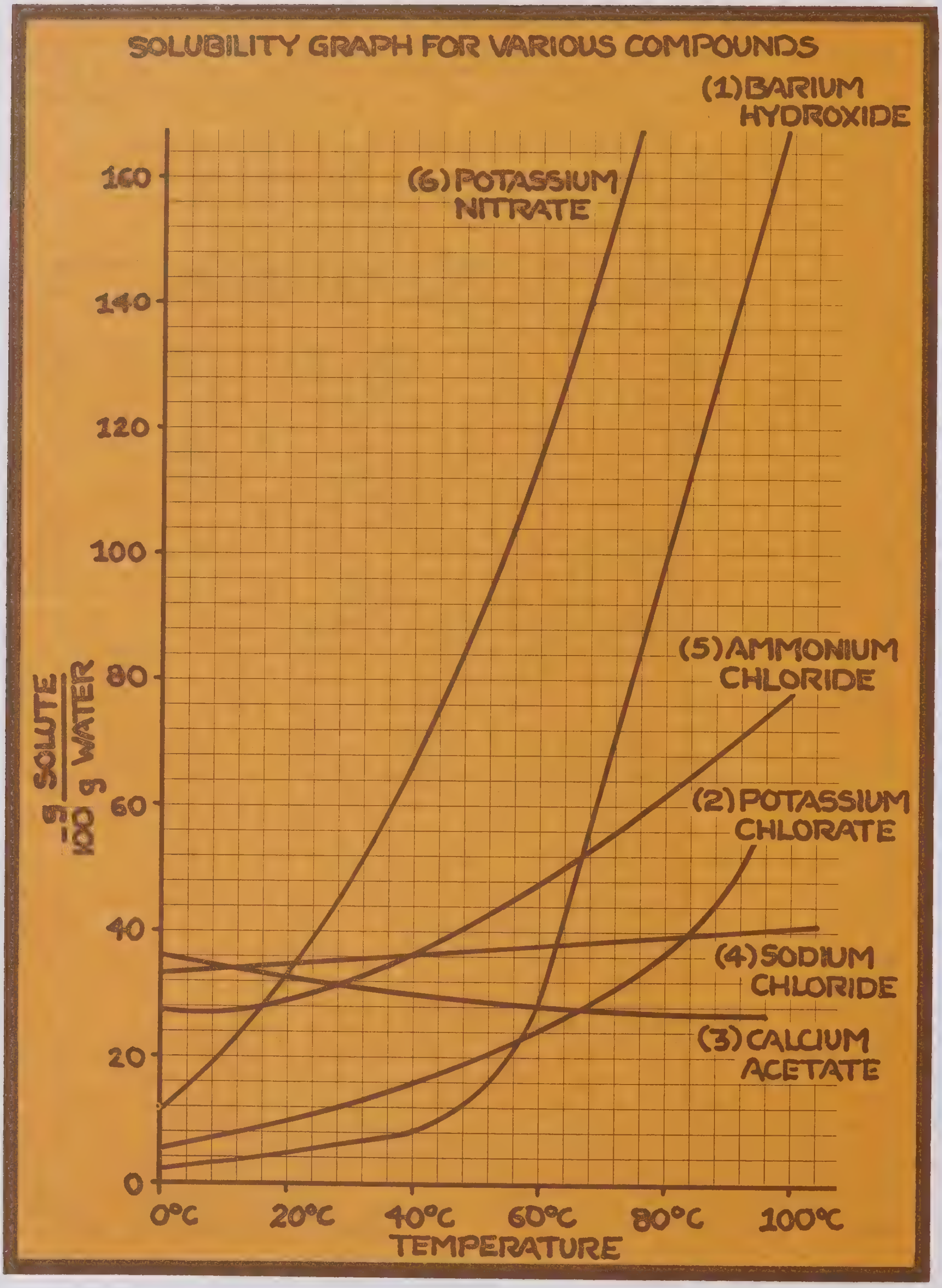
The curve that you have plotted is very similar to a curve used by chemists called a *solubility curve*. The difference between your curve and a solubility curve is fairly minor. A solubility curve is a graph of temperature against grams of solute per 100 g of solvent in a saturated solution.

Here is what the solubility curve for copper sulphate looks like. On the following page you will find a composite graph of several solubility curves for other substances. Use these curves to answer the following questions.



SOLUBILITY CURVE OF COPPER SULPHATE

SOLUBILITY	TEMP.
14.3	0
17.4	10
20.7	20
25.0	30
28.5	40
33.3	50
40.0	60
55.0	80
75.4	100


SOLUBILITY GRAPH FOR VARIOUS COMPOUNDS
(1) BARIUM HYDROXIDE
(6) POTASSIUM NITRATE
(5) AMMONIUM CHLORIDE
(2) POTASSIUM CHLORATE
(4) SODIUM CHLORIDE
(3) CALCIUM ACETATE
g SOLUTE / 100 g WATER
160
140
120
100
80
60
40
20
0
0°C
20°C
40°C
60°C
80°C
100°C
TEMPERATURE

Solubility Chart for Inorganic Compounds

$\frac{\text{grams Solute}}{\text{100 g Water}}$

Substance	*0°C*	*20°C*	*40°C*	*60°C*	*80°C*
barium nitrate	5.0	9.2	14.2	20.3	27.0
calcium chloride	59.5	74.5	102.0		
cupric chloride	41.4	43.5	45.6	47.7	49.8
ferric chloride	74.4	91.8		325.0	525.8
potassium bromide	53.5	65.2	75.5	85.5	95.0
potassium permanganate	2.83	6.4	12.56	22.2	
ammonium chloride	29.4	37.2	45.8	55.2	65.6
sodium chloride	35.7	36.0	36.6	37.3	38.4
lead nitrate	38.8	56.5	75.0	95.0	115.0
potassium chromate	58.2	61.7	65.2	68.6	73.0
potassium dichromate	5.0	13.1	29.2	50.5	73.0

Solubility Chart for Gases

$\frac{\text{grams Gas}}{\text{100 g Water}}$

Gas	*0°C*	*20°C*	*40°C*	*60°C*	*80°C*
carbon dioxide	0.05	0.11	0.17	0.26	
hydrogen	4.42	5.20	5.78	5.80	5.73
nitrogen	4.09	5.75	7.60	8.70	
oxygen	1.91	2.95	4.14	4.84	5.28
nitrous oxide	0.07	0.15	0.24		

Questions and Problems

1. What meaning would you attach to the following terms:
   (a) soluble;
   (b) slightly soluble;
   (c) very soluble;
   (d) concentrated;
   (e) dilute.
2. Comment on the statement: "Solutions are always transparent."
3. Distillation is a process that can be used to separate the constituents of a solution or some mechanical mixtures. Define the term. State the changes of state that are involved in this method of separation. Outline briefly why this method of separation will work. What property of the constituents is used?
4. Separation of a mechanical mixture or a solution is based upon a difference in the properties of the two constituents. State what property(s) could be used to separate the following mixtures:

(a) iron filings in sand;
(b) silver in finely crushed rock;
(c) oil in water;
(d) cream in milk;
(e) dust in air;
(f) tars in cigarette smoke;
(g) carbon monoxide in car exhaust.

5. Pick out the mechanical mixtures in Question 4. Describe a mechanical process that could be used to separate the constituents of these mixtures.
6. 10 g of potassium chloride are dissolved in water.
   (a) What information is conveyed to you by the above statement?
   (b) State the information about the solution that is *not* conveyed to you by the above statement.
   (c) Reword the statement to make it more meaningful.
7. How is a maximum amount of solute specified by the definition given for solubility?
8. When giving a solubility is it necessary to state the solvent used? Explain.
9. Why is temperature specified when giving a solubility?
10. General speaking, how does solubility vary with temperature?
11. What is the solubility of sodium chloride at 50°C?
12. What is the maximum amount of barium hydroxide that will dissolve into 100 g of water at 65°C?
13. What is the maximum amount of ammonium chloride that will dissolve in 25 g of water at 40°C?
14. If 15 g of potassium chlorate is dissolved in 100 g of water at 70°C, what type of solution will be formed? Explain.
15. What would happen to a saturated solution of copper sulphate at 60°C if it were cooled to 20°C? How much copper sulphate would precipitate out?
16. If you were given a saturated solution of potassium chlorate at 10°C which was then heated to 30°C, how much additional potassium chlorate could be dissolved?
17. Which of the substances for which you have data is most soluble at 50°C? Which is the least soluble?
18. If you were given a clear colourless solution of sodium thiosulphate (hypo) how would you test it to see if it was a saturated, unsaturated or supersaturated solution? Describe in some detail the test that you would use.
19. Use the data given below to plot a graph of solubility versus temperature.

Grams Solute per 100 g Solvent	
Solubility	*Temperature (°C)*
28	0
30	10
35	20
38	25
53	35
73	45
86	50
140	60
165	65

(a) What is the solubility of this substance at 30°C? 15°C?

(b) At what temperature will (i) 100 g (ii) 50 g of this substance just dissolve in 100 g of water?

(c) A 10 ml sample of the solution at 27.5°C was found to contain 5.5 g of the solute. What can you conclude about this sample?

(d) If 50 ml of a saturated solution of this substance were seeded and cooled from 60°C to 30°C, how much solute would you expect to precipitate out?

(e) If 25 ml of a saturated solution of this substance were heated from 15°C to 45°C, how much more solute could be dissolved at the higher temperature?

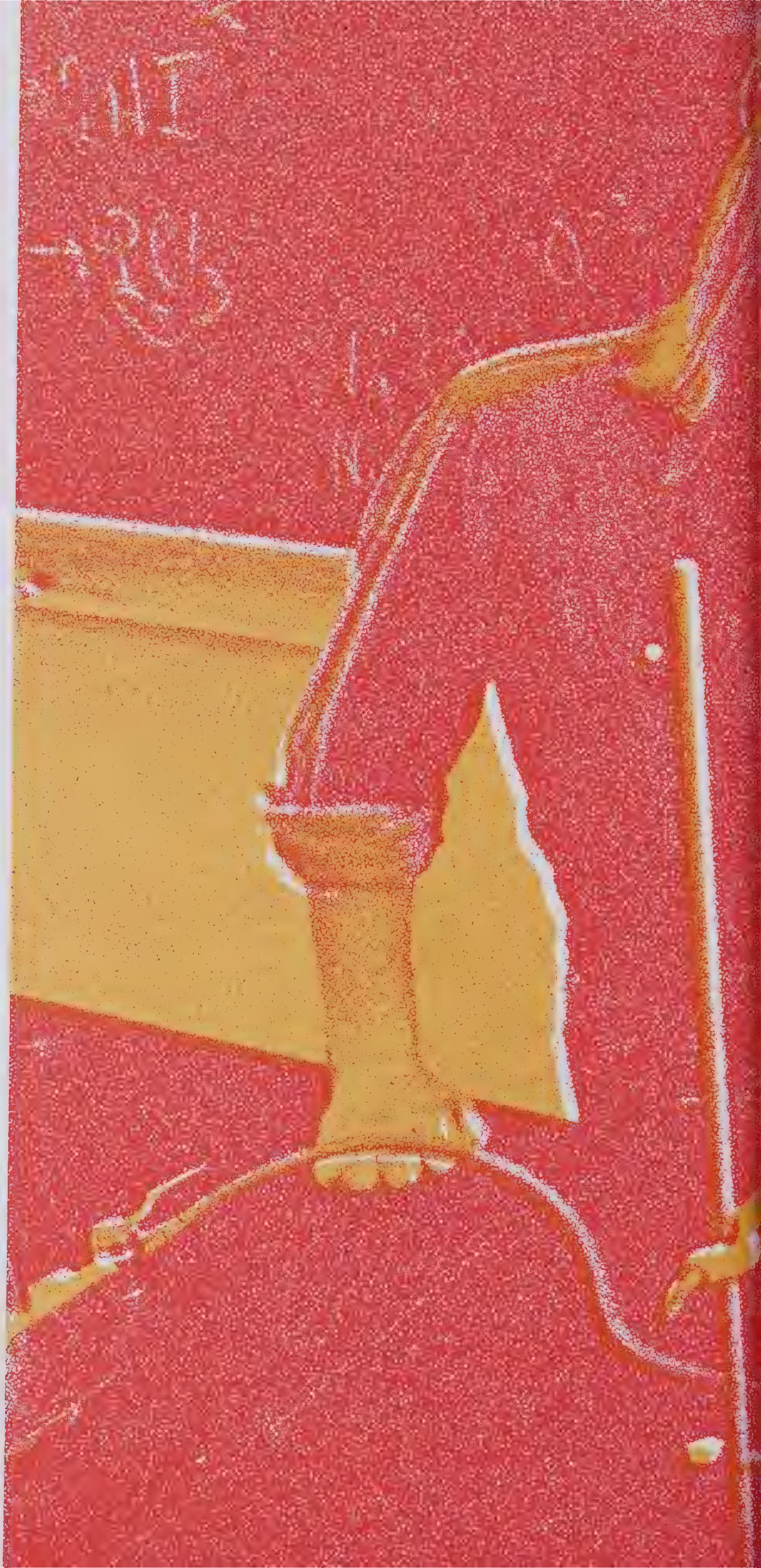
(f) How much of this substance could be dissolved in 750 g of water at 50°C?

Some Words You Should Know

- Observation
- Conclusion
- Phlogiston
- Theory
- Senses
- Qualitative
- Quantitative
- Regularity
- Conclusion
- Conditions that matter
- Interpretation
- Hypothesis
- Physical change
- Chemical change
- States of matter
- Calorimetry
- Solutions
- Mixtures
- Colloids
- Homogeneous
- Heterogeneous
- Phase
- Distillation
- Filtration
- Density
- Transparent
- Clear
- Coloured
- Solubility
- Solubility Curve

2

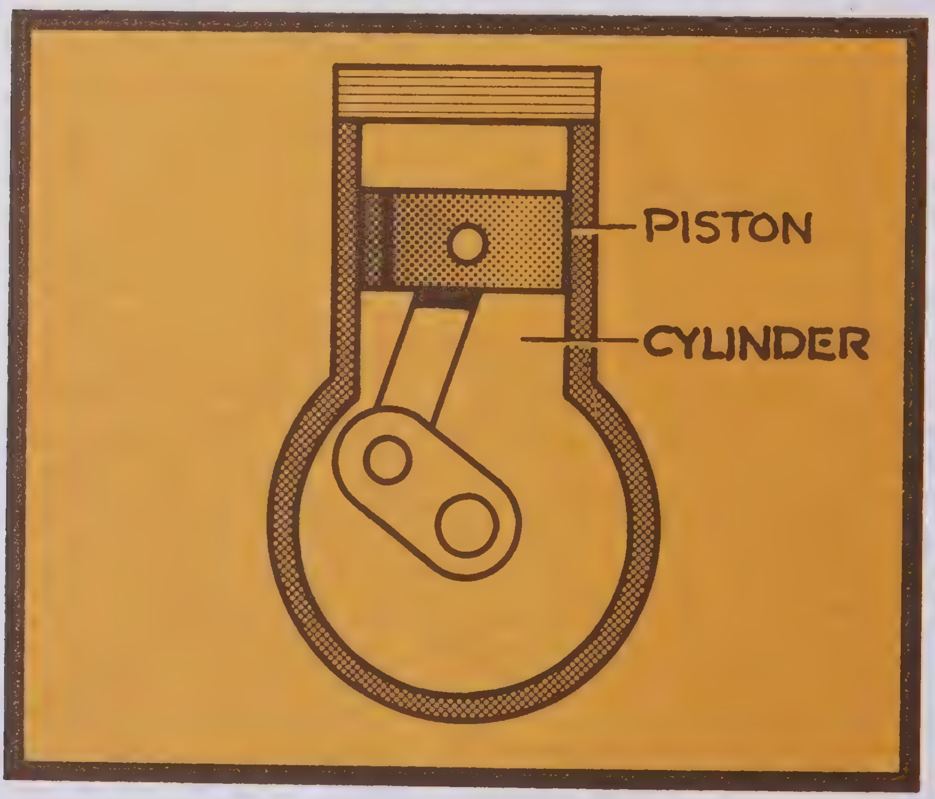
Conservation of Matter/Composition of Compounds/Decomposition of Compounds/Law of Constant Composition/Boyle's Law/Charles' Law/The Particle Model of Matter/Avogadro's Principle/Molar Volume/Ideal Gas Law/Partial Pressure/Vapour Pressure/Elements/Symbolism/The Mole



Phlogiston Theory

In the seventeenth century a number of unusual discoveries about air pressure were made. One of great significance involved the development of a vacuum and the work that could be done with air pressure. It was learned that a vacuum could be created by other means than pumping air out of a container. This involved heating water to the gaseous state then allowing the steam to condense into a liquid in a closed system. If one wall of the closed system were movable, like a piston in a cylinder, it was discovered that air pressure pushing against this movable wall could be used to create a vacuum. After the air pressure had pushed the wall in, heating the chamber again could push the wall out. This principle forms the basis of the steam engine and by 1700 such a steam engine had been developed. Numerous changes in design, especially by James Watt, led to a very workable steam engine toward the end of the eighteenth century.

Two things are very significant about this discovery. First, man was no longer dependent on his own muscles or those of animals to do work and, secondly, there was an increased interest in the use of fire to do different things. This new interest in fire directed chemists' interest toward combustibles. Questions arose about why some things burn readily while others do not, and what exactly burning or combustion is.



The Greeks thought that things that would burn contained the element, fire. In 1669 J. J. Becker proposed that solid materials contained three kinds of earth, one of which was the substance that would burn. George Ernst Stahl, following Becker's ideas, proposed what is known as the Phlogiston Theory. Stahl's proposal was that various materials contained phlogiston. Combustible materials had a high phlogiston content that was released to the air when the substance burned. To Stahl, air played only a very small part in the burning process and was of very little significance. Stahl further proposed that reactions like the tarnishing of silver or the rusting of iron were much like combustion reactions as phlogiston left the substances as they reacted. His theory met with little opposition and was widely accepted throughout the eighteenth century.

When wood, for example, burned, it lost mass because of the escaping phlogiston. A piece of metal rusting gained mass even though it was supposed to be *losing* phlogiston. Stahl's theory did not really resolve this dilemma. It may seem a dilemma to us, but to the eighteenth-century chemist masses and accurate measurement were not very important so the problem was not a serious one.

Modern chemists are just the opposite. Accuracy in measurement and in recording observations is extremely important. Keep both the phlogiston theory and accuracy in mind as you work through this book.

Changes in Matter

Already in our chemistry course we have observed that matter is capable of changing. For example, in Experiment 1.5 iodine crystals changed as did benzoic acid crystals. Copper mixed with sulphur changed markedly when the two were heated in combination. We concluded that these changes were either physical or chemical in nature and we established a basis for deciding on the type of change that occurred.

In Experiment 2.1 you will mix various substances together as you look for evidence of change in the original substances. It will help if before doing the experiment you make a table in your notebook that includes the following:

(a) description of the substances before mixing;
(b) observations during the reaction;
(c) description of substances after the reaction;
(d) how you know a change occurred.

2.1

Mixing Substances

ammonium chloride
sodium hydroxide
lead nitrate
potassium iodide
anhydrous copper sulphate
sodium carbonate
1 M hydrochloric acid

1. Record the temperature of 2 cm of water in a test tube. To the water in the test tube add about 5 cm^3 of ammonium chloride. Shake the test tube. Record the temperature again.
2. Repeat Step 1, replacing the ammonium chloride with two pellets of sodium hydroxide.

Warning

Sodium hydroxide must not come in contact with your hands. Wash the resulting solution down the drain.

3. Dissolve a few crystals of lead nitrate in 2 cm of distilled water. Also dissolve a few crystals of potassium iodide in the same amount of distilled water. Describe each solution. Mix the two solutions together.
4. Dissolve a few crystals of anhydrous copper sulphate in 2 cm of water.
5. To 2 cm of dilute hydrochloric acid in a test tube add 5 cm^3 of sodium carbonate.

Questions

1. The substances that dissolved in the water seemed to disappear. How do you account for this?
2. What reactions were:
   (a) exothermic?
   (b) endothermic?
3. Outline the various ways that a change in the substances was indicated in this experiment.
4. How do you account for the change observed in each part of the experiment?
5. Generally, each experiment that we do generates new questions that we do not have answers for. Make a list of the questions that this experiment has raised in your mind.

In all five parts of Experiment 2.1, a change occurred in each of the substances that were mixed. If we think back to the basis we established for distinguishing chemical from physical change, we will realize that the changes that occurred in Experiment 2.1 were chemical in nature. As we will encounter numerous chemical reactions during this course, we should investigate the laws that govern these reactions.

Laws Governing Reactions

We will begin by looking at what happens to mass in a chemical reaction. As usual, we shall first experiment to gather some facts, then we shall look for the regularity(s) and attempt to make some generalizations.

2.2

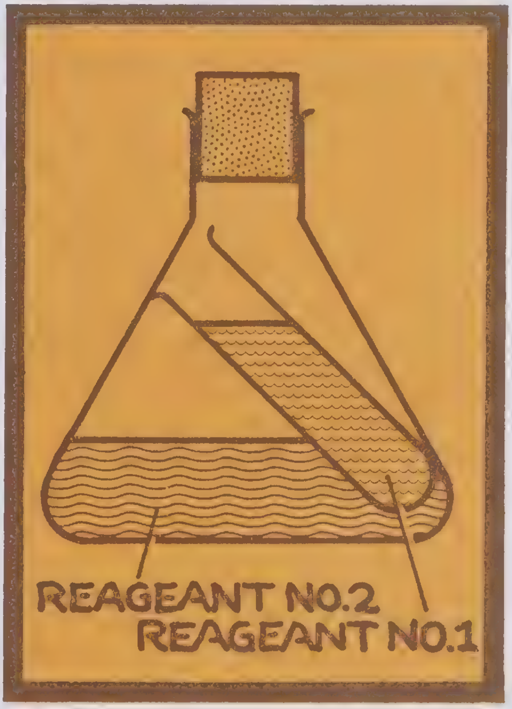
Mixing Solutions

Each of you will be assigned a pair of reagents from the following list. Each of the reagents is already dissolved in water to form a fairly dilute solution.

1. (a) lead nitrate
   (b) potassium iodide
2. (a) potassium chlorate
   (b) lead nitrate
3. (a) barium chloride
   (b) sodium sulphate
4. (a) silver nitrate
   (b) potassium chloride
5. (a) iron nitrate
   (b) potassium thiocyanate

You will also need:

250 ml erlenmeyer flask
small test tube
stopper
balance



1. Find a test tube that will stand as shown in the 250 ml erlenmeyer flask.
2. Put about 20 ml of one reagent in the flask.
3. Put 20 ml of the second reagent in the test tube. Carefully describe both reagents.
4. Cautiously lower the test tube into the flask. Be very careful not to spill the solution from the test tube into the flask.
5. Stopper the flask firmly.
6. Carefully determine the mass of the flask, stopper and contents.
7. As soon as you have recorded the mass, invert the flask several times to mix the reagents and immediately determine the mass of flask, stopper and contents again. Record this mass.
8. Describe the contents of the flask.

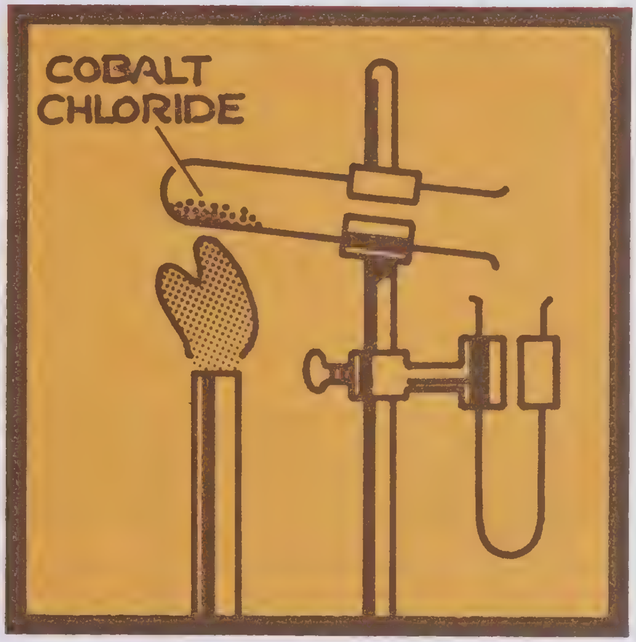
Questions

1. Was there a chemical reaction inside the flask? Explain.
2. How does the mass of the flask and contents after the reaction compare to the mass of the flask and contents before the reaction?
3. How do your results compare with other groups who used the same pair of reagents?
4. How do your results compare with groups who used different reagents?
5. Can you find a regularity here that relates changes of mass to chemical reactions?
6. Do you think there is any generalization that might be made from this very limited set of results?
7. Could you explain why you got the results you did? What does this experiment tell you about the nature of matter?

The nature of matter has long intrigued men of science. There seemed to be so many different kinds of matter in a variety of forms. Scientists for years have known that heat has a considerable effect on many different forms of matter. In your experiments, heat lighted the candle or changed the benzoic acid or the copper sulphate. In Experiment 2.3 you will heat samples of different substances, taking particular notice of the mass of the substances.

2.3 Heating Substances

retort stand
2 clamps
2 test tubes
bunsen burner
balance
glass plate
cobalt chloride



Part A

1. Determine the mass of an empty test tube.
2. Place about 1 cm of cobalt chloride into the test tube. Determine the mass of the tube and its contents.
3. Clamp this test tube and a second one as shown in the diagram.
4. Heat the cobalt chloride until it becomes a uniform colour throughout. Record your observations as you heat the test tube.
5. Allow the test tube to cool. Determine the mass of the tube and contents again. Remove the contents of the test tube and place them on a glass plate. Divide the solid into two parts.
6. Pour a few drops of distilled water on one part. Put a few drops of the liquid from the vertical test tube on the other part.
7. Draw a table like Table 2.1 in your notebook and complete it.

Table 2.1

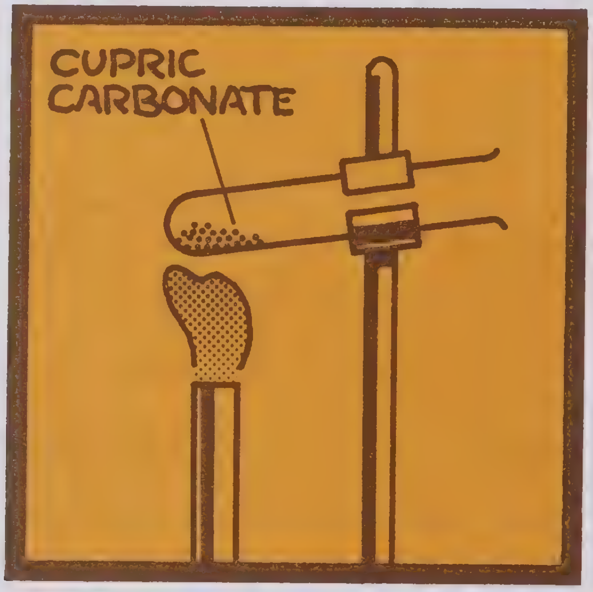
Mass of test tube empty ______________ g
Mass of test tube and cobalt chloride before heating ______________ g
Mass of cobalt chloride ______________ g
Mass of test tube and contents after heating ______________ g
Mass of contents ______________ g

Questions

1. What happened to the cobalt chloride when it was heated?
2. What does this tell you about the nature of cobalt chloride?
3. What is the liquid in the vertical test tube? What observation(s) would lead you to this conclusion?
4. What happened to the solid residue when water was added to it?
5. What is a substance like cobalt chloride called?
6. What does this tell you about the nature of matter?

Part B

test tube
retort stand and clamp
bunsen burner
balance
cupric carbonate



1. Determine the mass of a clean dry test tube.
2. Put about 1.5 cm of cupric carbonate into this test tube. Determine the mass again.
3. Clamp the tube as shown in the diagram and heat with a hot bunsen flame. Continue to do so until the colour is uniform throughout.
4. Allow the tube to cool. Determine the mass of the tube and contents again. Save the contents for Part C.
5. Draw a table like Table 2.2 in your notebook and complete it.

Table 2.2

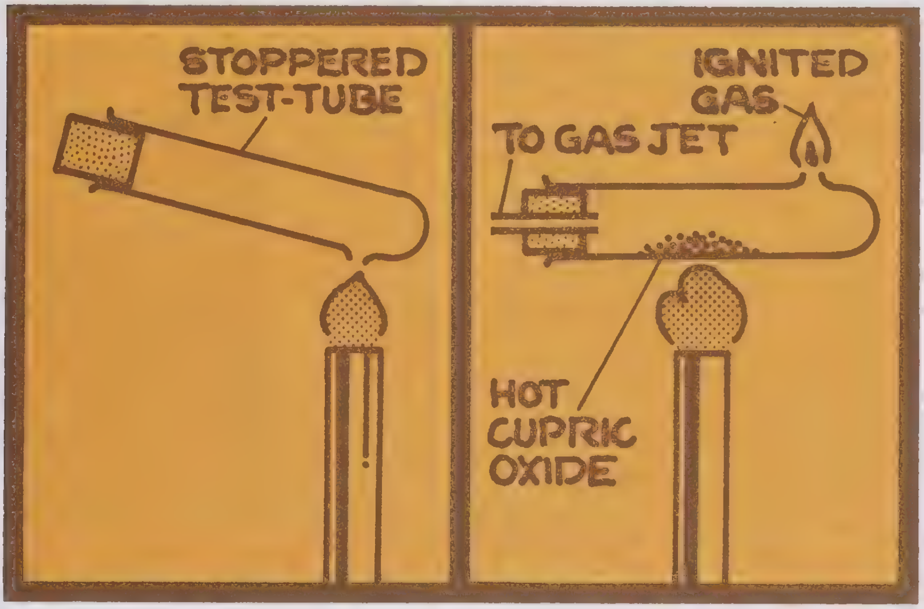
Mass of empty test tube	______ g
Mass of test tube and cupric carbonate	______ g
Mass of cupric carbonate	______ g
Mass of test tube and contents after heating	______ g
Mass of contents	______ g

Questions

1. What evidence was there that the cupric carbonate changed to a new substance?
2. What in your opinion happened to the cupric carbonate?
3. What does this tell you about the nature of matter?

Part C

residue from Part B
stopper
test tube with a hole (see diagram)
bunsen burner
one-holed stopper
glass tube



1. Place a stopper into an 18 × 150 mm test tube. Heat the lower side of the tube as shown in the diagram until a hole appears in the side. Remove the stopper and allow the tube to cool.
2. Determine the mass of the test tube that you have just punctured.
3. Transfer the residue from Part B to this test tube and determine the mass again.
4. Assemble the appartus as shown in the diagram.
5. Pass a gentle stream of tap gas through the test tube and ignite it at the hole you have made. Heat the test tube and contents strongly with your bunsen burner until an obvious change has occurred in the substance. Turn off the bunsen burner, then the gas tap to the test tube. Remove the hose from the gas tap and allow the tube to cool.
6. Determine the mass of the test tube again. Examine the contents.
7. Draw a table like Table 2.3 in your notebook and complete it.

Table 2.3

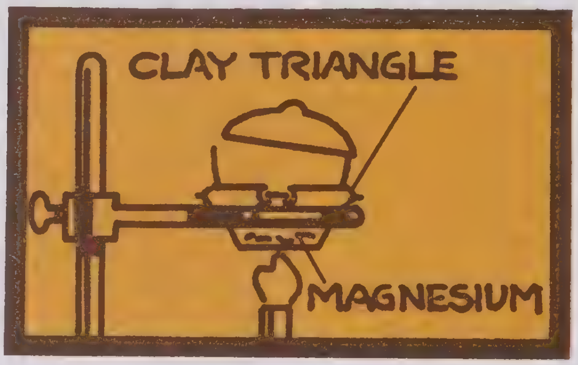
Mass of tube and black residue	______________ g
Mass of tube	______________ g
Mass of black residue	______________ g
Mass of tube and brown residue	______________ g
Mass of brown residue	______________ g

Questions

1. What evidence was there that the black residue (copper oxide) changed to a new substance?
2. What in your opinion happened to the copper oxide?
3. What does this tell you about the nature of matter?

Part D

crucible and lid
retort stand and ring
clay triangle
bunsen burner
balance
magnesium ribbon



1. Determine the mass of a clean dry crucible and lid.
2. Measure off 10 cm of magnesium ribbon. Clean the ribbon by pulling a piece of folded sandpaper over it once or twice.
3. Wind the magnesium ribbon around a pencil to form a spiral. Place the spiral in the crucible and determine the mass of the crucible, lid and magnesium.
4. Support the crucible on a clay triangle on a retort ring as shown in the diagram.
5. Place the lid on the crucible, leaving a small opening for air to enter.
6. Heat the crucible with a bunsen burner adjusted to give a very hot flame for 5 to 10 min.
7. Cautiously slide the lid open a bit. If the contents start to burn with a bright white flame replace the lid and heat for another 3 to 5 min. Check again and if necessary continue heating. When the magnesium no longer flares up, remove the lid carefully and heat for a minute, then allow the crucible and lid to cool.
8. When the crucible, lid and contents are cool enough determine their mass again.

9. Arrange your data in a table like Table 2.4. Calculate the mass of the magnesium, the mass of the product and the difference between them.

Table 2.4

Mass of crucible, lid and product	________ g
Mass of crucible and lid	________ g
Mass of product	________ g
Mass of crucible, lid and magnesium	________ g
Mass of crucible and lid	________ g
Mass of magnesium	________ g
Difference in mass between product and reagent	________ g

Questions

1. How do you account for the change in mass in this experiment?
2. What substance may be responsible for this change in mass?
3. What does this tell you about the nature of matter?

Calculations

Part A

Complete this statement from your observation chart.
In ________ g of cobalt chloride (amount you used) there are ________ g water (loss of mass). Therefore in 100 g of cobalt chloride there are ________ g of water.

Part B

Complete this statement.
In ________ g of cupric carbonate there are ________ g of carbon dioxide (loss of mass). Therefore in 100 g of cupric carbonate there are ________ g of carbon dioxide.

Part C

Complete this statement.
In ________ g of copper oxide there are ________ g copper (residue left). Therefore in 100 g of copper oxide there are ________ g of copper.

Part D

Complete this statement.

In ____________ g of magnesium, ____________ g of oxygen react. Therefore with 100 g of magnesium ____________ g of oxygen react.

Compare class results for each part of the experiment by completing a table on the blackboard like Table 2.5.

Table 2.5

Part A

Mass of Cobalt Chloride Used	*Grams of Water in 100 g*	

Part B

Mass of Cupric Carbonate Used	*Grams of Carbon Dioxide in 100 g*	

Part C

Mass of Cupric Oxide Used	*Mass of Copper in 100 g*	

Part D

Mass of Magnesium Used	*Mass of Oxygen with 100 g*	

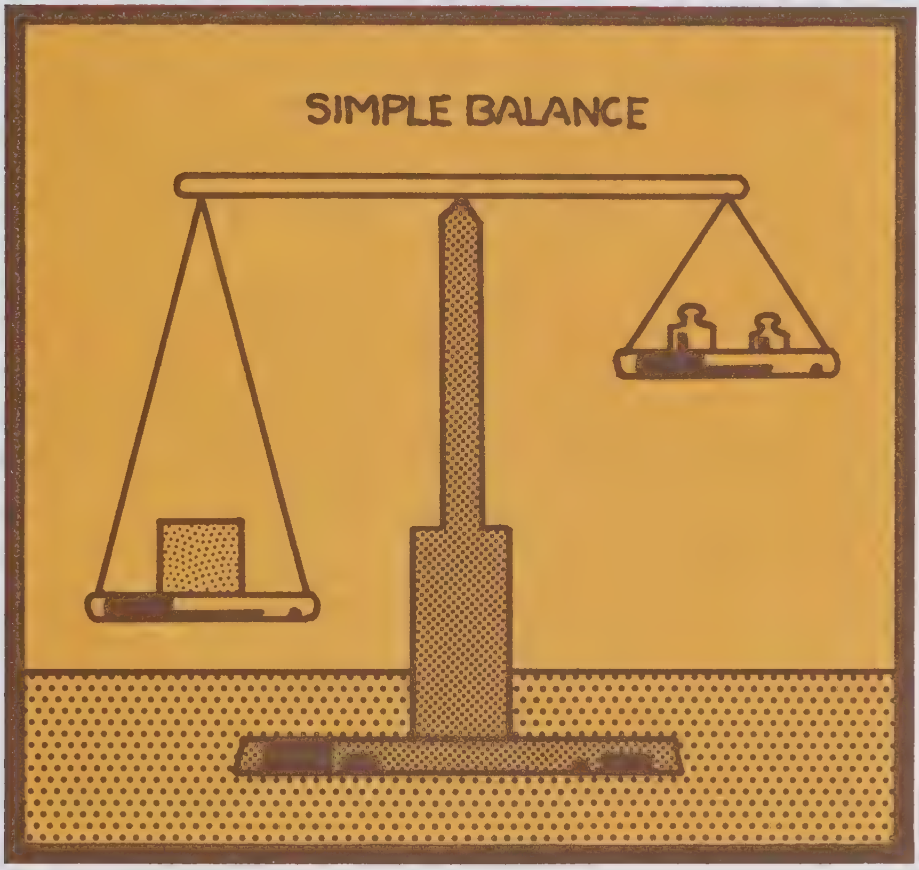
Analysis and Questions

1. How does the amount of cobalt chloride used in each case compare? Were they the same?
2. How does the amount (g) of water in 100 g of cobalt chloride compare in each case?
3. What do these results suggest about cobalt chloride?
4. If cobalt chloride is characteristic of compounds in general, then what statement could we make about the composition of compounds?
5. Answer Questions 1, 2 and 3 for the other three substances used

in Parts B, C and D. Do the results obtained here support the statement that you made in Question 4? Explain.

6. Does this series of experiments tell us anything about the nature of matter?

Measurement and Chemical Change



Accurate measurement was of little importance to the chemist of the eighteenth century. What really mattered was whether he could explain changes in appearance and properties of substances. A small change in mass as a substance underwent a chemical change could be ignored.

As the industrial age developed, however, this thinking changed and a growing interest in precision and accuracy developed. More precise equipment such as balances and thermometers were produced. This resulted in a greater concern for changes in mass which had previously been ignored. Such apparatus led to experiments like Experiment 2.2, which you recently completed. In that experiment your observations of a low solubility substance as each solution was mixed indicated that a chemical reaction had occurred, yet your measurements indicated that the mass remained constant. Since there was not an increase or decrease in mass during the chemical change, then one is led to the conclusion that matter is neither created nor destroyed in the course of a chemical reaction. When this hypothesis was first put forward, thousands of experiments were carried out in which the mass of substances was determined before and after reaction. This hypothesis has stood the test of time and is now regarded as a law, the Law of Conservation of Mass.

If Experiment 2.3 was carried out with a certain degree of accuracy you will have already recognized a second regularity that also leads to a chemical law.

Consider for example the heating of the cobalt chloride. A student who selected a 4 g sample should have found that 2.77 g of water were released upon heating. His statement would be: in 4 g of cobalt chloride there are 2.77 g of water. Therefore in 100 g of cobalt chloride there are

$$\frac{100}{4} \times 2.77 \text{ g of water}$$

$$= 7 \text{ x } 10 \text{ g of water}$$

Another student with a 2 g sample of cobalt chloride would state: in 2 g of cobalt chloride there are 1.38 g of water. Therefore in 100 g of cobalt chloride there are

$$\frac{100}{2} \times 1.38 \text{ g of water}$$

$$= 7 \text{ x } 10 \text{ g of water} .$$

All the other students, whether they had 4 g or 2 g or any other mass of cobalt chloride in their sample would also find that there was 7 x 10 g of water in each 100 g sample. The conclusion, of course, is that every 100 g sample of cobalt chloride contains the same amount of water or, more generally, every sample of cobalt chloride has a constant amount of water.

This hypothesis is now a chemical law as it too has withstood the test of many experiments. It is known as the Law of Constant Composition or Definite Proportions. This law states that every pure sample of the same chemical compound contains the same substances combined in the same proportions by mass.

Like all good scientists, we do not base such conclusions on one set of facts. This is why you did Parts B, C and D. If we allow for small differences in experimental error, the Law of Constant Composition should be rendered true by these additional experiments.

Problems

1. Mercuric oxide is a bright orange compound that decomposes readily when heated. Describe in detail how this substance could be used by a class to verify the Law of Constant Composition.

2. (a) 1 g of copper was converted to copper oxide by first reacting the copper with nitric acid to produce copper nitrate. The copper nitrate was then heated and yielded 1.25 g of copper oxide.
   (b) Copper hydroxide was converted to copper oxide and water by heating. The 3 g sample of copper oxide produced was then converted to copper and water with hydrogen. 2.39 g of copper were produced.
   (c) A sample of copper carbonate was converted to copper oxide by heat. The oxide was then converted to copper by heating with tap gas. A 1.5 g sample of copper oxide released 0.30 g of oxygen as the copper was formed.

   Show how these three experiments support the Law of Constant Composition.

The Nature of Matter

Chemistry, by definition, is a study of the composition, structure and transformations of matter. We have already found that matter exists in at least three different forms: solid, liquid or gas. Further, by either adding or removing heat, matter can be transformed from one form to another. An ice cube, for example, may melt and become a liquid which may vaporize into the gaseous state. Each change is brought about by adding heat. Removing heat may reverse the process. In addition, matter in its various forms may react chemically with other kinds of matter. Copper reacts with sulphur, magnesium reacts with oxygen, and lead nitrate reacts with potassium iodide. Our most important discovery to date, however, has been that these changes are governed by rigid laws. In such changes, mass remains constant even though a new substance with a new set of properties is produced. And finally, when new substances are formed the materials making up these substances always combine in some fixed or definite proportion by mass.

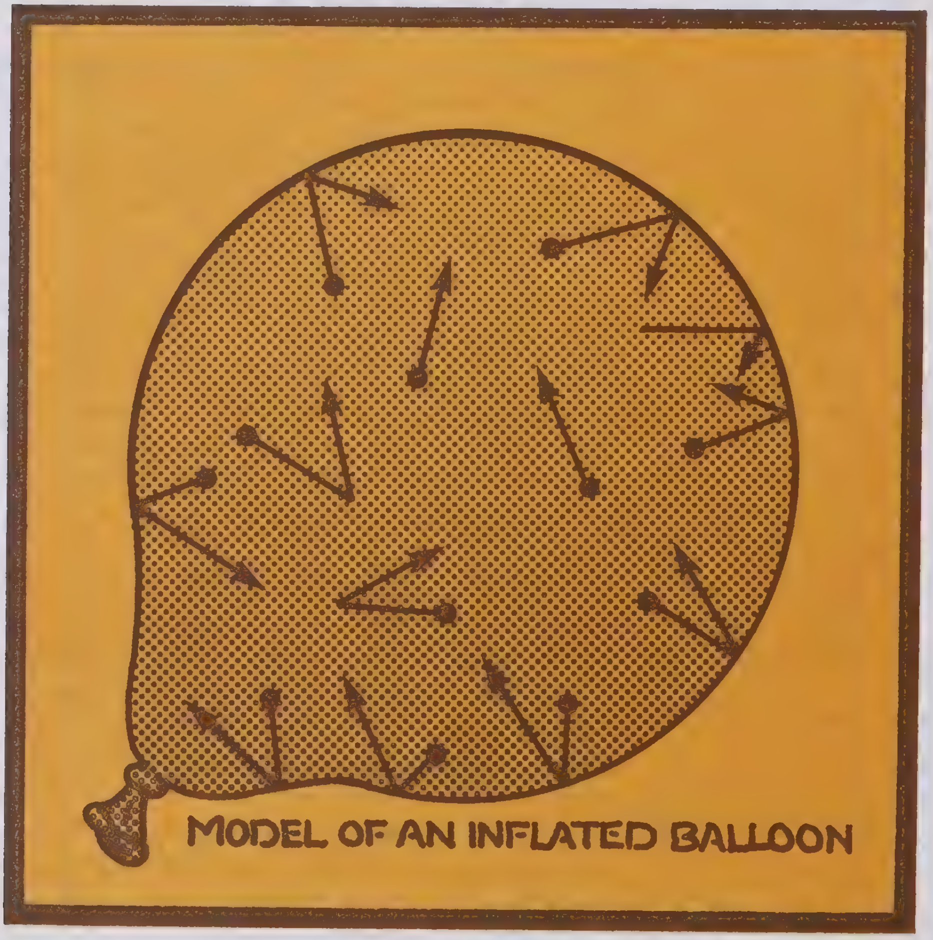
The obvious questions that come to mind must surely be, how is matter made up so that these things can happen? How could a lump of iodine become an almost colourless vapour that floats in the air but will still reunite into a solid lump again? How could two clear colourless solutions produce a bright yellow solid? Why should the mass not change when such a solid is formed? Why should compounds always consist of substances combined in a definite proportion by mass? Why should a kettle of water boil away leaving nothing visible behind? The questions could continue. Perhaps you have unanswered questions of your own that have been generated by the experiments you have done. The task before us now is to delve into the unknown in an attempt to discover something about the actual makeup of matter.

Our studies will now be directed toward substances in the gaseous state even though they may seem elusive or mysterious. The alchemists must also have felt this way about gases and vapours since liquids that vaporized easily were called "spirits," a word carrying an obvious sense of mystery.

Consider for a moment what happens when you fill a balloon with air from your lungs. The balloon obviously gets larger and the larger it gets the harder you have to blow because the air that you have put in it is pushing back at you. If the balloon is tied, it remains a constant size or if punctured it bursts without any visible sign of what gave it the initial properties. What is really the nature of the air that we blew into the balloon?

Two assumptions are going to be presented here. Based on the sublimation of iodine or benzoic acid, we will assume that matter is made up of small particles. Secondly, since an inflated balloon does not immediately collapse, we will assume that when such particles collide with each other or with the walls of the balloon no energy is lost and the particles are perfectly elastic.

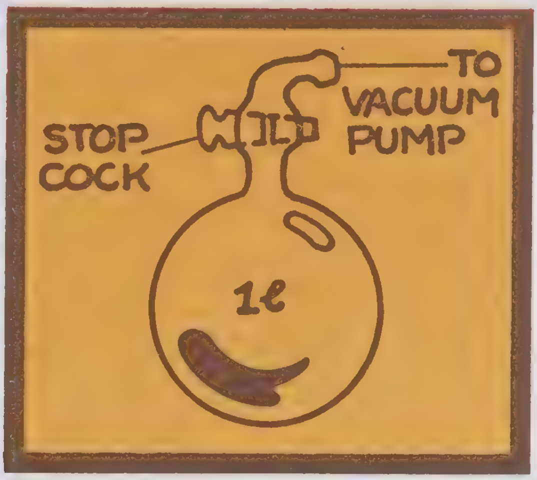
Working on this hypothesis, try these experiments with gases.


MODEL OF AN INFLATED BALLOON

2.4

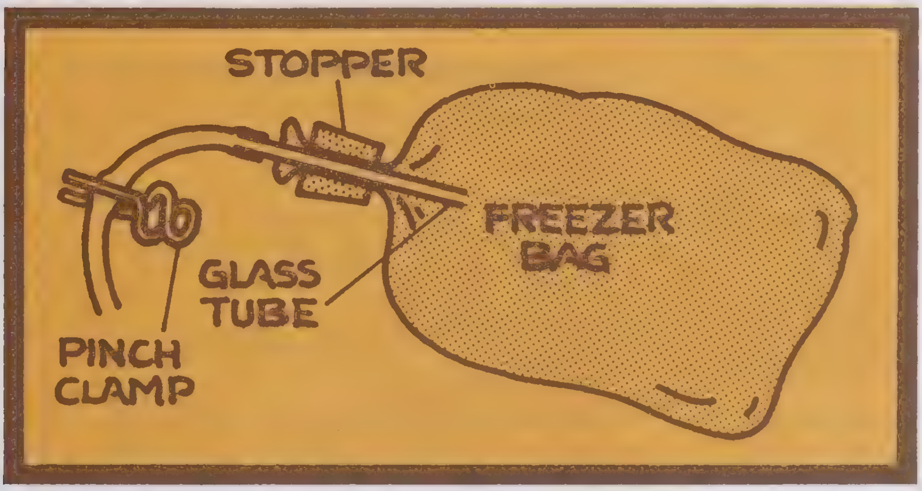
Determining the Mass of Gases

gas bag assembly
balance
large jug
delivery tube
1000 ml graduated cylinder
various gases
density of air flask



Part A

1. Remove all the air from a gas bag by rolling it into a tight roll. Clamp the inlet tube.
2. Carefully determine the mass of the empty gas bag while it is clamped. The mass should be within 0.01 g if at all possible.
3. Blow air into the bag until it becomes firm. Hold the bag by the stopper in a vertical position and open the clamp quickly to equalize pressure in the bag with the atmospheric pressure.
4. Determine the mass of the assembly again.
5. Determine the mass of a density of air flask full of air and again after it has been evacuated.



Questions

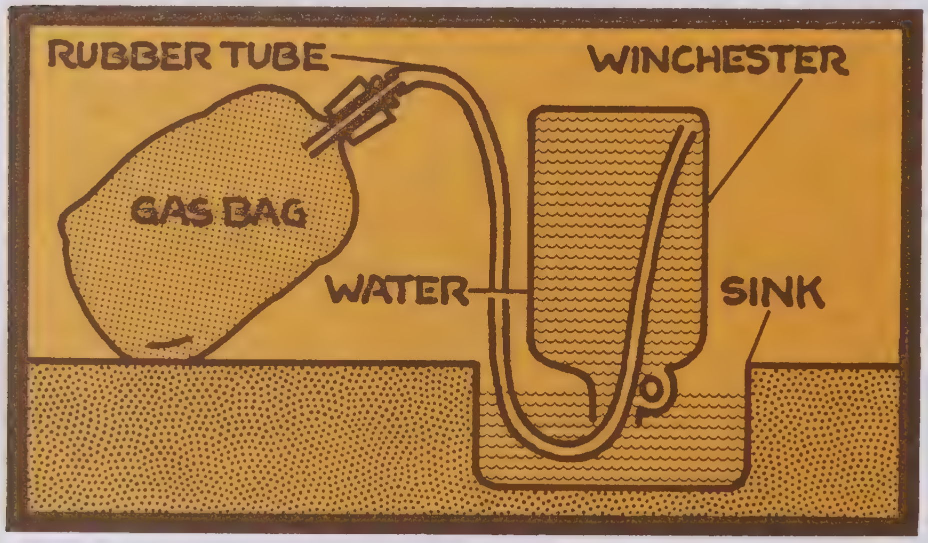
1. Allow 0.02 g for experimental error. Determine the apparent mass of air in the bag.
2. Determine the mass of air in the density flask.
3. Account for the apparent mass of air in the bag.

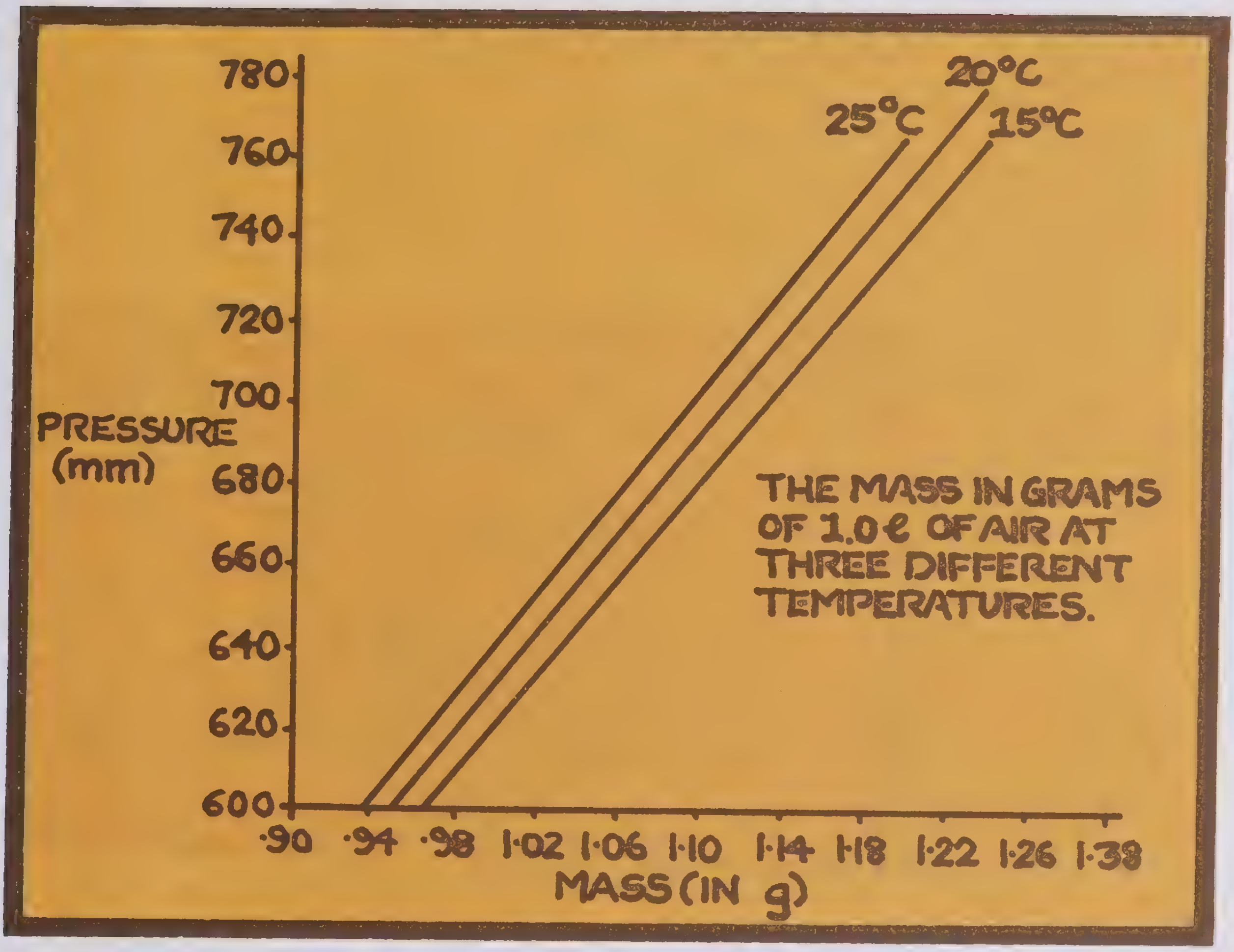
Part B

1. If the gas bag assembly is not the same one used in Part A, then you must begin once again by determining the mass of the assembly when it is completely free of gas.
2. Fill the bag in turn with each of the gases provided by your teacher. Determine the mass of the bag and gas after filling it.

Remember to equalize pressure with the atmosphere each time before determining the mass.

3. After you have determined the mass of each bag of gas, find the volume of the bag by displacement of water in a large jug. This is done as follows.
   (a) Fill the jug with water and determine the total volume of water in the jug by pouring the water into a large graduated cylinder.
   (b) Fill the jug again and invert it in a sink partially filled with water. Run a delivery tube from the gas bag assembly, filled with air or some other gas, to the jug. The tube must reach right to the bottom of the jug. Remove the clamp and force the gas from the bag into the jug. When the bag is empty remove the delivery tube, stopper the jug and turn it upright. Measure the volume of water remaining in the jug.
4. Record the atmospheric pressure and the room temperature.





Calculations

1. Calculate the volume of the bag. This is the volume of the jug minus the volume of water remaining in the jug.
2. Using the graph provided, calculate the mass of air displaced by the bag.

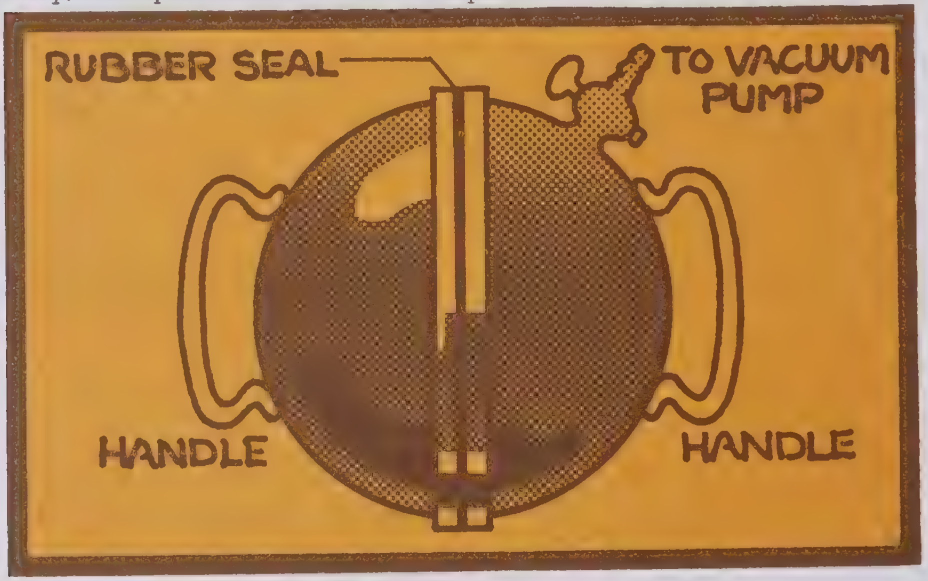
This mass represents the loss of mass due to buoyancy.

3. Determine the apparent mass of each of the gases used.
4. Determine the actual mass of each of the gases used. This is the apparent mass plus the mass of air displaced.
5. Assuming the particle hypothesis is true, how could you explain the different masses for equal volumes of different gases. Remember that there may be more than one explanation and you should try to give all possible explanations.

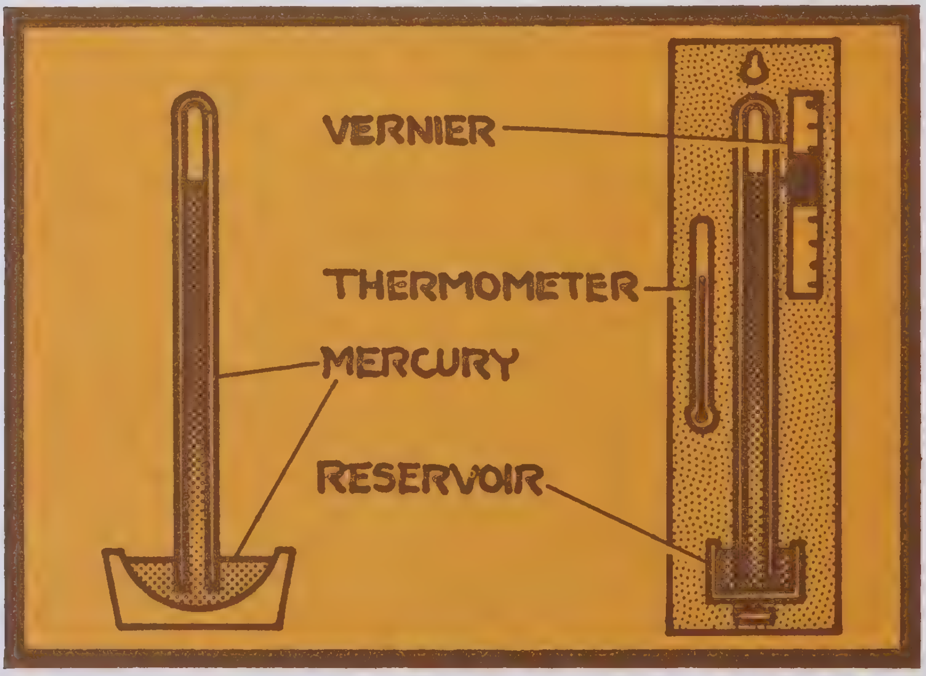
In order to decide which explanation best fits the facts we will need more information about gases. Our explanations so far refer to both pressure and temperature. Perhaps we should look into these a bit more.

Gas Pressure

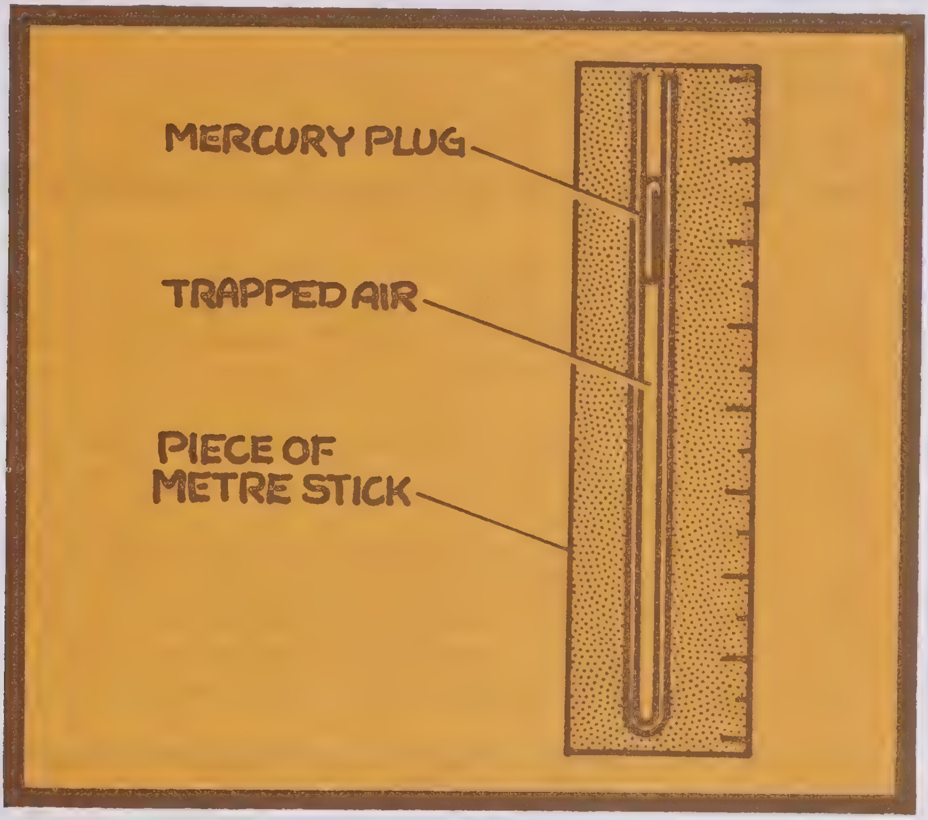
A German physicist, Otto van Guericke, illustrated that gases, especially air since it is the most common gas, were capable of exerting tremendous pressures. Guericke, with the aid of an air pump, drew the air from two specially designed hemispheres. Teams of horses were attached to each hemisphere and pull as they may, the sphere would not fall apart.


RUBBER SEAL
TO VACUUM PUMP
HANDLE
HANDLE

Even before Guericke, an Italian physicist had shown that air could exert pressure and further showed that it was capable of supporting a mercury column enclosed in a glass tube. This experiment led to the development of the mercury barometer.

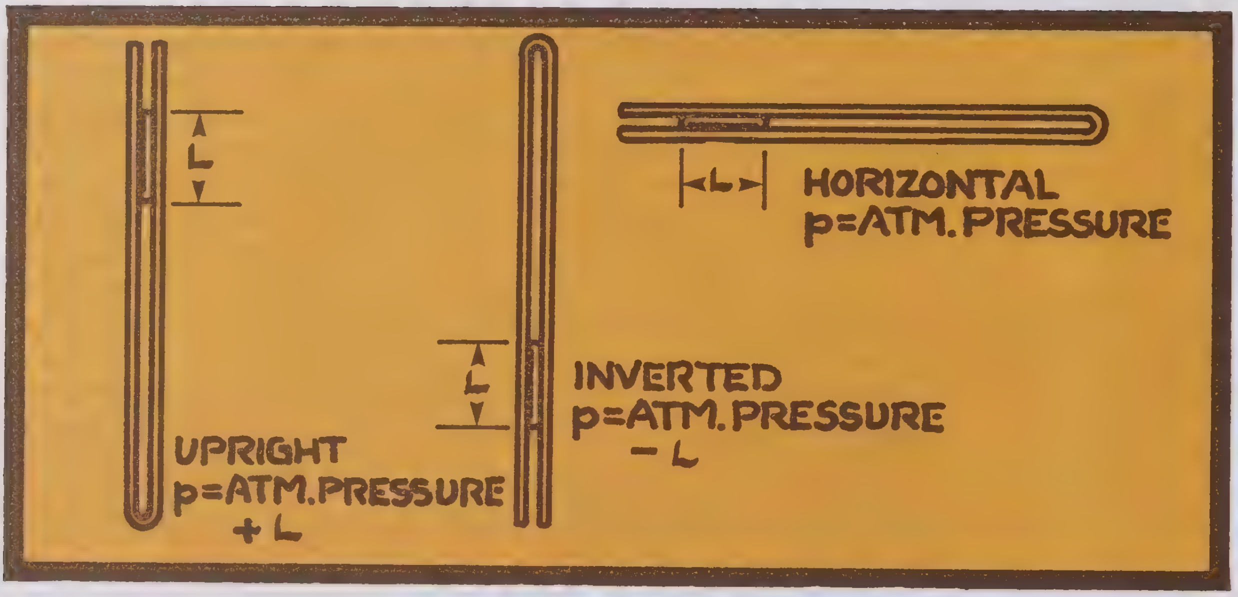

VERNIER
THERMOMETER
MERCURY
RESERVOIR

A simple mercury barometer can be made with a dish of mercury and a thick-walled glass tube about 80 cm in length closed off at one end. Mercury from the dish is placed in the tube with an eye dropper until the tube is completely filled. The open end is then covered and the tube is inverted in the dish of mercury (see diagram). The mercury falls part way out of the tube, but not completely. The resulting column is about 76 cm in length with an evacuated space at the top of the tube. Air pressure is holding the mercury up in the tube. If the air pressure were to increase the mercury column would increase in length and a decrease in pressure would result in a decrease in the length of the mercury column.



Air pressure, therefore, is capable of supporting a mercury column about 76 cm in height.

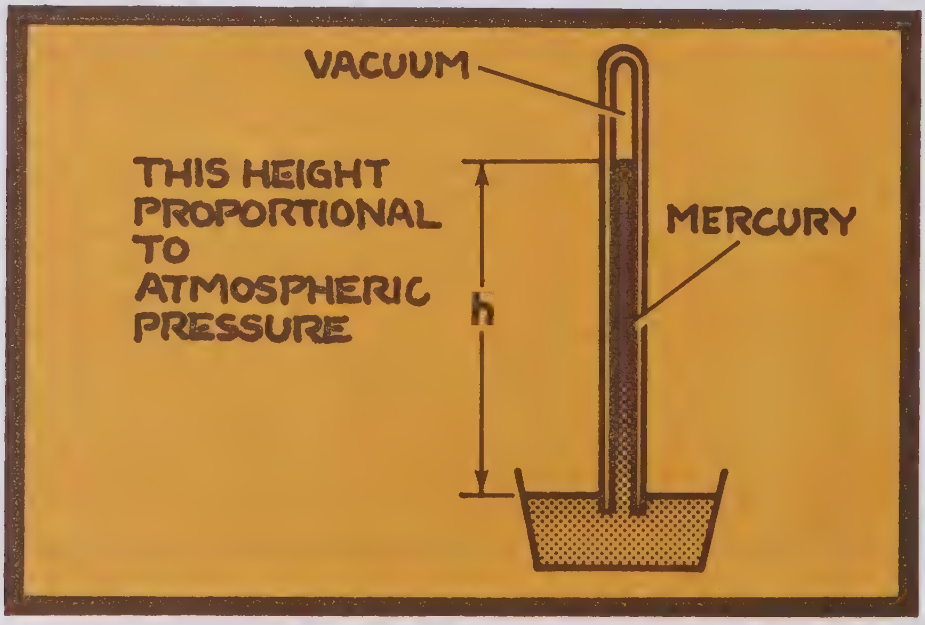
Consider for a moment the apparatus shown in the diagram. It contains some trapped air and a mercury plug. When the tube is held in a vertical position the trapped air is pushed on by the air pressure (which is equivalent to a column of mercury about 76 cm in length) plus the mass of the mercury plug. In an inverted position, part of the air pressure is used just to hold the mercury plug in the tube. The pressure on the trapped air will, therefore, be air pressure minus the length of the mercury plug. Finally, in a horizontal position, the mercury plug tends not to fall out nor does it exert any pressure on the trapped air. The pressure on the trapped air in a horizontal position will just be the air pressure with no influence from the mercury plug.


L
HORIZONTAL
p=ATM. PRESSURE
L
INVERTED
p=ATM. PRESSURE
− L
L
UPRIGHT
p=ATM. PRESSURE
+ L

Position	*Pressure of Trapped Air*
Upright	Air Pressure (mm Hg) + Mercury Plug (mm)
Horizontal	Air Pressure (mm Hg)
Inverted	Air Pressure (mm Hg) − Mercury Plug (mm)

This means that by holding the apparatus in any one of the three positions a different pressure may be exerted on the confined air (gas).

Pressure is defined as a force per unit area. In the old English system, a common pressure unit was pounds per square inch (PSI). In SI, the unit of pressure is called the pascal (Pa), the unit of area is the square metre (m^2) and the unit of force is called the newton (N). The pascal is equal to a newton per square metre ($1 \text{ Pa} = 1 \text{ N/m}^2$).


VACUUM
THIS HEIGHT PROPORTIONAL TO ATMOSPHERIC PRESSURE
h
MERCURY

In our next experiment we will need to know the pressure exerted by the atmosphere on a sample of a gas. The instrument used to measure this pressure is the barometer. Under average conditions the pressure exerted by the atmosphere (atmospheric pressure) is very nearly 100 kPa (more accurately 101.3 kPa). This pressure is large enough to support a column of mercury 760 mm high in a mercury barometer. As the pressure falls, the level of the mercury falls. As the pressure rises, the level of the mercury rises. The longer the mercury column, the greater the atmospheric pressure. The length of the mercury column is therefore proportional to the atmospheric pressure so that:

$$760 \text{ mm Hg} = 101.3 \text{ kPa}.$$

In the next experiment we will measure the pressure by measuring the lengths of the mercury columns involved. Since the pressure is proportional to the length of the mercury columns, we will record the pressures in millimetres of mercury. We will not convert to pascals, the accepted SI unit of pressure.

2.5

Pressure Change and Gas Volume

apparatus illustrated on page 92

Warning

This apparatus must be handled with extreme care and the mercury plug must be in one piece.

1. Record room temperature and the air pressure at the start of this experiment.
2. Hold the apparatus in a horizontal position. This may best be done by laying it flat on your desk. Measure the length of the mercury plug as well as the length of the air column. *Accuracy is extremely important*. Make your measurements in millimetres.
3. Repeat Step 2 holding the tube in an upright vertical position
4. Repeat Step 2 holding the tube in an inverted vertical position.
5. Record room temperature and air pressure now.

Calculations

The air column is in the shape of a cylinder ($V = \pi r^2 h$).

1. Calculate the volume of air in the tube in each of the three positions.

2. Calculate the pressure on the trapped air in the tube in each of the three positions.

Complete a table in your notebook similar to Table 2.6 for each tube used.

Table 2.6

Position	*Mercury Plug (mm)*	*Air Column (mm)*	*Volume Air (V)*	*Pressure (p)*
Horizontal *Upright* *Vertical*				

Questions

1. How does (a) the mass and (b) the temperature of the confined gas compare for each position? Explain.
2. Compare values for the product $p \times V$ for any particular tube. Do you recognize any regularity? Make a generalization based on this regularity. Relate the product $p \times V$ at constant temperature to the quantity of gas in the system by a mathematical statement.
3. Explain the qualitative relationship between pressure and volume in terms of particles; that is, explain why the volume decreases as the pressure increases on a given mass of gas.

The chart below gives the results of an experiment much like the one that you have just completed.

Pressure	*Volume*	$p \times V$
$p_1 = 760$ mm Hg	$V_1 = 10$ ml	$p_1 \times V_1 = 7600$
$p_2 = 691$ mm	$V_2 = 11$ ml	$p_2 \times V_2 = 7600$
$p_3 = 844$ mm	$V_3 = 9$ ml	$p_3 \times V_3 = 7600$

(Products are approximate)

We may therefore say that $p_1 \times V_1 = p_2 \times V_2 = p_3 \times V_3 =$ constant, or simply: $p_1 \times V_1 = p_2 \times V_2$. p_1 and V_1 represent the original pressure and volume while V_2 represents the new volume when p_1 is changed to p_2.

$$\therefore p_1V_1 = p_2V_2$$

$$\text{or } V_2 = \frac{p_1 \times V_1}{p_2}$$

This mathematical expression tells us that when the pressure changes, the new volume is equal to the original volume times a fraction made up of the new and the original pressures. Consider the following problems.

Example 1

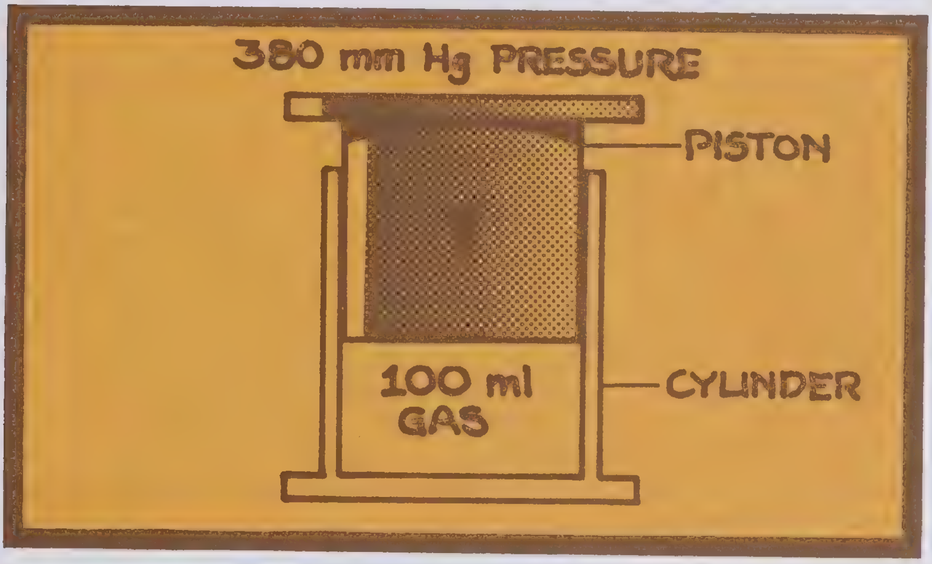
100 ml of air are trapped in a cylinder with a pressure on the gas equivalent to 380 mm Hg. The pressure is suddenly increased to 760 mm. What does the volume become?

1. The pressure increases so the volume decreases.
2. New volume = original volume (100 ml) × a fraction made up of the two pressures that will give a reduction in volume.

$$\therefore \text{pressure fraction} = \frac{380 \text{ mm}}{760 \text{ mm}}$$

$$\text{The new volume is} = 100 \text{ ml} \times \frac{380 \text{ mm}}{760 \text{ mm}}$$

$$= 50 \text{ ml}$$



Example 2

If the pressure was reduced from 380 mm to 190 mm what would the volume become?

1. The pressure decreases so the volume increases.
2. New volume = original volume (100 ml) × pressure fraction that will give an increase in volume.

$$\therefore \text{pressure fraction} = \frac{380 \text{ mm}}{190 \text{ mm}}$$

$$\therefore V = 100 \text{ ml} \times \frac{380 \text{ mm}}{190 \text{ mm}}$$

$$= 200 \text{ ml}$$

Problems

1. A 1 ℓ sample of oxygen at 25°C and 760 mm pressure has the pressure changed to 700 mm. If the temperature remains at 25°C, what will the volume become?

2. A cylinder of compressed air contains 5 ℓ of air at 20°C and 500 kPa pressure. What volume would this air occupy at 100 kPa pressure?
3. 600 ml of hydrogen gas are collected by a student during an experiment. The pressure on the hydrogen is 750 mm. What volume of hydrogen would have been collected if the pressure had been 760 mm? (Assume that temperature remains constant.)

Robert Boyle

Robert Boyle was born to wealthy Irish parents on January 26, 1627. Boyle was extremely clever and his parents could afford to give him a good education. He studied French, Latin, Greek and Hebrew.

While still only a young boy, he entered Eton College, a very distinguished preparatory school. Like any wealthy English gentleman, Boyle was removed from school to tour the European countries. While visiting Italy when he was only fourteen years old, he encountered Galileo. This distinguished man had such an influence on Boyle that Boyle decided to study science as a lifetime career.

Boyle entered Oxford University to study science. He found himself with a highly influential group of scholars that eventually became the Royal Society.

Boyle is a name familiar to any first-year chemistry student in high school because of Boyle's Law, a mathematical equation that tells how a gas behaves under changes in pressure when the temperature is kept constant.

The experiment that led to the famous Boyle's Law stemmed from a study of the mercury barometer. He designed a tube in the shape of a "U" but with one side of the "U" considerably shorter than the other and closed at the shorter end. Mercury was added to fill the curve of the tube trapping some air into the closed end. The mercury level was adjusted so that it was the same height in each arm of the tube. This meant that the pressure on the confined gas was equal to atmospheric pressure. He then added enough mercury through the open end of the tube to exactly double the pressure on the confined gas. The volume of the gas was reduced to exactly half the original value. When the pressure was increased to three times the original value the volume of the gas was reduced to one third the original volume. Boyle expressed the relationship in a mathematical expression that remains today as Boyle's Law.

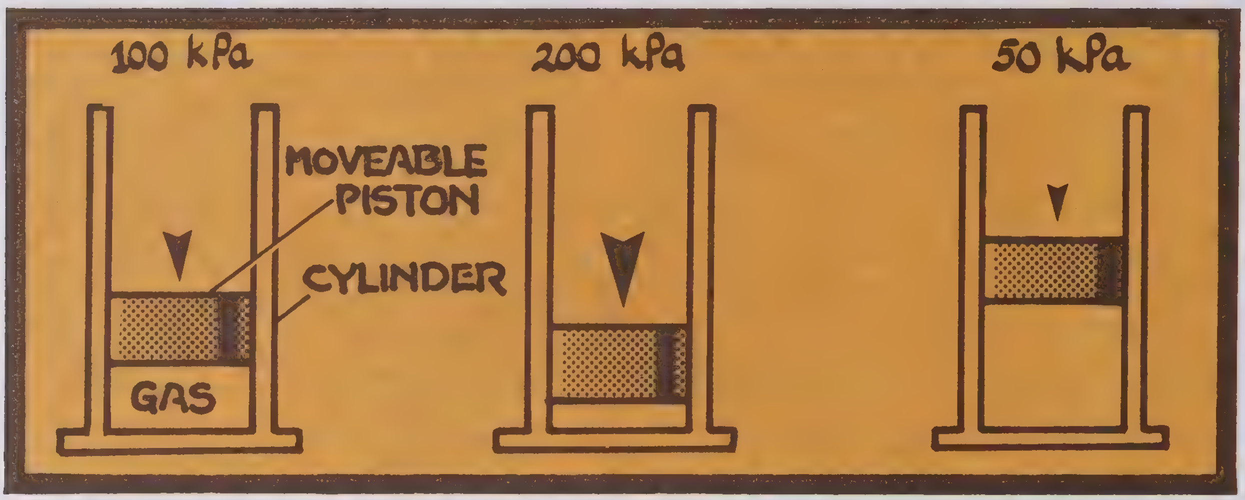
Boyle investigated a number of different areas of science including sound, electricity and colour. Many of his achievements were recorded in letters written to a cousin.

Boyle died at the age of 64 on December 30, 1691.

Robert Boyle (1627-1691)

Gases and Pressure — Temperature Change

Robert Boyle had been attempting to explain the various heights of mercury in a mercury barometer when he arrived at the relationship illustrated in Experiment 2.5. This relationship was presented to scientific circles and to the public in 1662. Boyle's discovery may be stated as: When the pressure on a given mass of gas is increased the volume will decrease. The relationship is, of course, only true if the temperature is kept constant. What this statement tells us is that if the pressure on 100 ml of a gas at constant temperature is increased from 100 kPa to 200 kPa the volume will be reduced to 50 ml, or if the pressure on the 100 ml of gas is changed from 100 kPa to 50 kPa the volume will go to 200 ml.



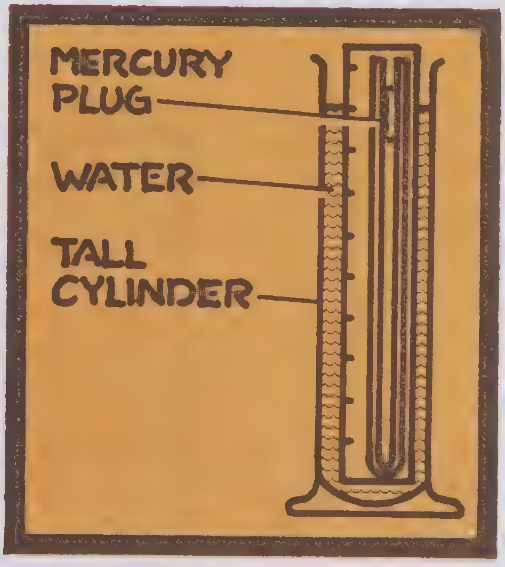
The significance of Boyle's discovery should not be underestimated. For years engineers have taken advantage of this relationship in designing equipment that involves the compression of gases. One of the most common, of course, is the piston-driven engine of cars, motor bikes or lawn mowers. By knowing the volume of gas being taken into the cylinder and by what amount it will be compressed, it is possible to determine by a simple mathematical calculation just how great the pressure will be. The equipment can then be designed accordingly.

In the next experiment you will use the same piece of apparatus as you have just used. However, this time the pressure must remain *constant*. How can this be done?

2.6

Temperature Change and Gas Volume

gas law apparatus
250 ml cylinder
thermometer



Keep the tube in the upright position.

1. Record the atmospheric pressure.
2. Add cold water and ice to a 250 ml graduated cylinder. Immerse the tube in the water-ice mixture until the entire air column is beneath the surface of the water. Place a thermometer in the cylinder. Allow the complete assembly to stand for 2 min so that the enclosed air will reach the temperature of the water.
3. Record the temperature of the water.
4. Record the length of the air column.
5. Repeat the same procedure at least four more times. Each time, however, the water should be about 15°C warmer than the previous time.
6. Record your observations and results in a table like Table 2.7.

Table 2.7

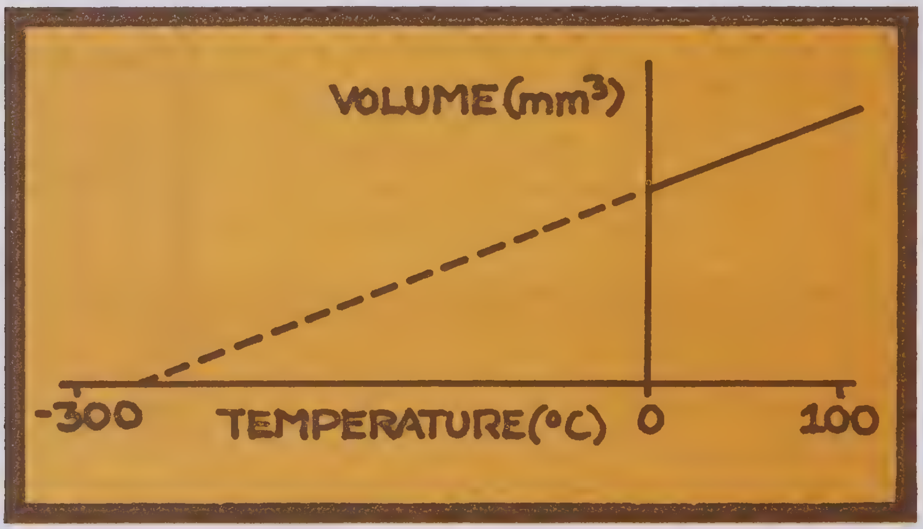
Tube # ______ *Atmospheric Pressure* ______

Temperature (°C)	Height of air column (mm)	*Volume of air* ($\pi r^2 h$)
1.		
2.		
3.		
4.		
5.		

Questions

1. Compare the pressure on the confined air for each part of the experiment.

2. Compare the mass of the confined air for each part of the experiment.
3. Plot your results on a graph. Set up the graph like the one shown below. Draw the best straight line through the points. By extrapolation from your graph, determine the temperature at which the line crosses the temperature axis.



4. Determine the equation corresponding to the straight line you have drawn. Use the form y = mx + b.

y = the volume of the gas V_g
m = the slope = b/a
x = the temperature of the gas t (°C)
b = the volume axis intercept = ma
a = the temperature axis intercept

Substituting in the original equation gives:

$V_g = mt + b$

But $b = ma$
$\therefore V_g = mt + ma$
$V_g = m(t + a)$
If we let $T = t + a$
then $V_g = mT$
or $V_g \propto T$ (m is a constant)

T is a new temperature scale based on the behaviour of gases. This temperature scale is called the kelvin scale.

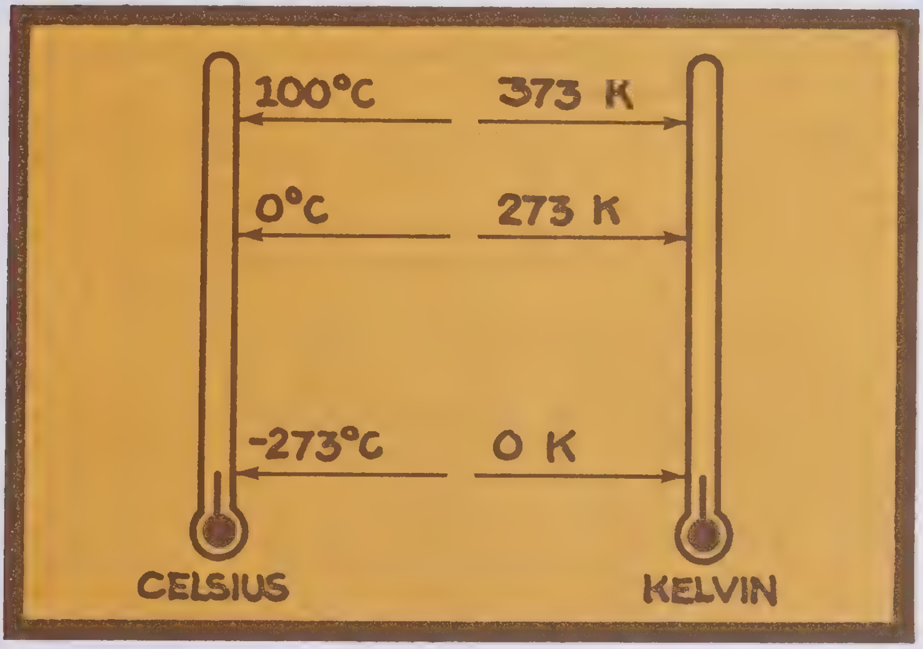
5. State the relationship between the volume of gas (V) and the temperature (t) in your own words. Include in your statement the conditions under which the statement holds true. This is called Charles' Law.
6. Explain the qualitative relationship between temperature and volume of a given amount of gas using your theory about gases.

Gases and Temperature Change

In Experiment 2.6 we found that the volume of a given mass of gas (pressure remaining constant) is altered as the temperature changes. Qualitatively you found that as the temperature increases, so does the volume and vice versa. But the shape of the graph indicated that this was not a direct relationship; that is, as the temperature (in °C) doubles the volume does not double. The volume changes directly with the kelvin temperature scale. On such a temperature scale zero corresponds to − 273°C.

kelvin temperature = degrees Celsius + 273

$$K = {}^\circ C + 273$$

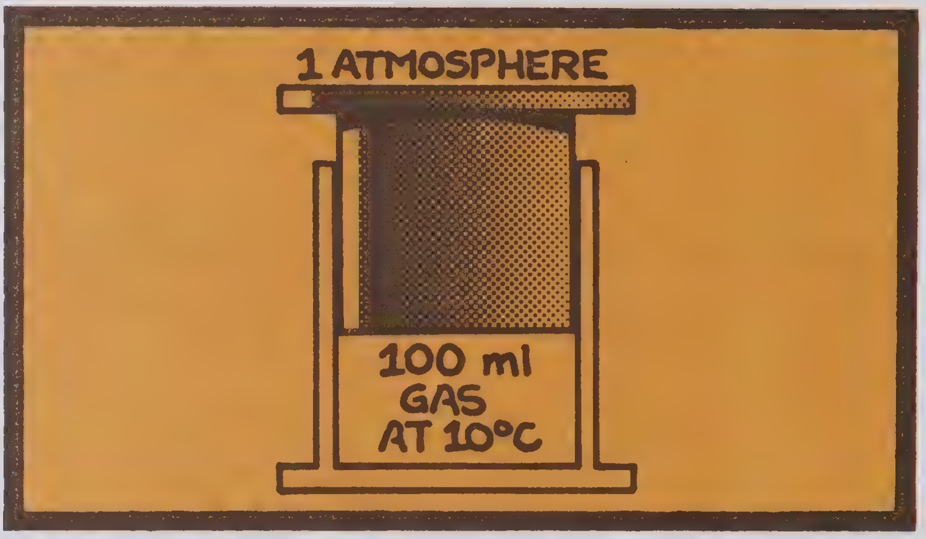


Once again, the significance of the relationship between volume and temperature should not be underestimated. Just as the pressure is altered in an internal combustion engine by a reduction in volume, the temperature will also have a tremendous influence and the engineer designing the engine must take this into account. You have probably noticed warnings on aerosol cans that state you must keep the can out of direct heat. The designer of that can had to consider the effects of heat when designing the can.

Consider the following problems.

Example 1

The cylinder in the diagram contains 100 ml of gas at 10°C. What will the volume become if the temperature is changed to 27°C?



1. As the temperature increases, the volume increases.
2. Volume varies directly with the kelvin temperature scale.

$$10^\circ\text{C} = 273 + 10 = 283\ \text{K}$$
$$27^\circ\text{C} = 273 + 27 = 300\ \text{K}$$

3. New volume = original volume × temperature fraction that will give an increase in volume $\left(\frac{300\ \text{K}}{283\ \text{K}}\right)$.

$$\therefore V = 100\ \text{ml} \times \frac{300\ \text{K}}{273\ \text{K}}$$
$$= 109.9\ \text{ml} = 110\ \text{ml}$$

Example 2

If the temperature of the 100 ml of gas was reduced to 0°C, what would the volume become?

1. As the temperature decreases, the volume decreases.
2. Volume varies directly with the kelvin temperature scale.

$$10^\circ\text{C} = 273 + 10 = 283\ \text{K}$$
$$0^\circ\text{C} = 273 + 0 = 273\ \text{K}$$

3. New volume = original volume × temperature fraction that will give a decrease in volume $\left(\frac{273\ \text{K}}{283\ \text{K}}\right)$.

$$\therefore V = 100\ \text{ml} \times \frac{273\ \text{K}}{283\ \text{K}}$$
$$= 96.5\ \text{ml}$$

Problems

1. In an experiment a student collected 3 ℓ of hydrogen at 18°C. If

the pressure remains constant what volume of gas will be present at 27°C?

2. A man at a concession booth at a fair is selling helium-filled balloons. Each balloon is filled in the morning with 5 ℓ of the gas. The morning temperature is 15°C but by mid-afternoon the temperatures has risen to 35°C. What volume of gas will be in each balloon now? (Assume pressure remains constant.)
3. A syringe containing 10 ml of air at 27°C is placed in the freezing compartment of a refrigerator where the temperature is − 13°C. What does the volume of air in the syringe become at this new temperature? (Assume pressure remains constant.)

Standard Temperature and Pressure

To make some reasonable comparison between the amounts of gas in various samples, both temperature and pressure must be noted. Further, some basic or standard temperature and pressure must be noted. Some basic standard temperature and pressure is selected and the volumes are determined at these standard conditions. Standard temperature is 0°C, while standard pressure is set at 760 mm Hg or 100 kPa. The need for this can easily be shown in an example.

The purchasing agent for a welding firm calls two suppliers of

Compressed gas cylinders. Large volumes of gases may be compressed and stored in cylinders such as these. A 10 ℓ cylinder of a compressed gas such as hydrogen can deliver as much as 1000 ℓ of gas at room temperatures. Acetylene, oxygen, hydrogen, nitrogen and carbon dioxide are but a few of the gases sold in this manner. *Courtesy of Canadian Liquid Air, Ltd.*

oxygen for price quotes. One is given at 10 ¢/ℓ at 15°C and 2500 kPa pressure, while the second is given at 8 ¢/ℓ at 27°C and 1500 kPa pressure. Wanting to get the best deal for his company, the agent decides on the 8 ¢/ℓ price. Does he make a wise choice? How can we find out?

In each case we will find out what volume of oxygen will be delivered at standard temperature and pressure (STP).

Quote 1

10 ¢/ℓ measured at 15°C and 2500 kPa pressure.

Volume at STP	0°C and 100 kPa pressure
Temperature	1. As the temperature decreases, the volume decreases. 2. 0°C = 273 K 15°C = 288 K The temperature fraction to give a decrease in volume is $\frac{273\text{ K}}{288\text{ K}}$
Pressure	1. As the pressure decreases, the volume increases. 2. The pressure fraction to give an increase in volume is $\frac{2500\text{ kPa}}{100\text{ kPa}}$ ∴ volume of oxygen delivered at STP is $1\ \ell \times \frac{273\text{ K} \times 2500\text{ kPa}}{288\text{ K} \times 100\text{ kPa}}$ $= 23.7\ \ell$ ∴ the cost of oxygen per litre delivered at STP is $\frac{10\text{¢}}{23.7\ell} = 0.4\text{ ¢}/\ell$

Quote 2

8 ¢/ℓ measured at 27°C and 1500 kPa pressure.

Volume at STP	0°C and 100 kPa pressure
Temperature	1. As the temperature decreases, the volume decreases. 2. 0°C = 273 K 27°C = 300 K The temperature fraction to give a decrease in volume is $\frac{273\text{ K}}{300\text{ K}}$
Pressure	1. As the pressure decreases, the volume increases.

2. The pressure fraction to give an increase in volume is

$$\frac{1500\ \text{kPa}}{100\ \text{kPa}}$$

∴volume of oxygen delivered at STP is

$$1\ \ell \times \frac{273\ \text{K} \times 1500\ \text{kPa}}{300\ \text{K} \times 100\ \text{kPa}}$$

$$= 13.8\ \ell$$

∴ the cost of oxygen per litre delivered at STP is

$$\frac{8¢}{13.8\ell} = 0.6\ ¢/\ell$$

This means that the purchaser made an unwise choice. He selected the more expensive supplier simply by not using some standard for a basis of comparison. The standard that chemists use is 0°C and 760 mm (100 kPa) pressure. This is referred to as *standard temperature and pressure*, often abbreviated as STP.

A further point to note in the example is that both the temperature and the pressure changed. To find the effect of these two changes simply consider the effects of one change on the volume and then the other. It does not matter which of the two factors you consider first.

Combined Gas Law Problems

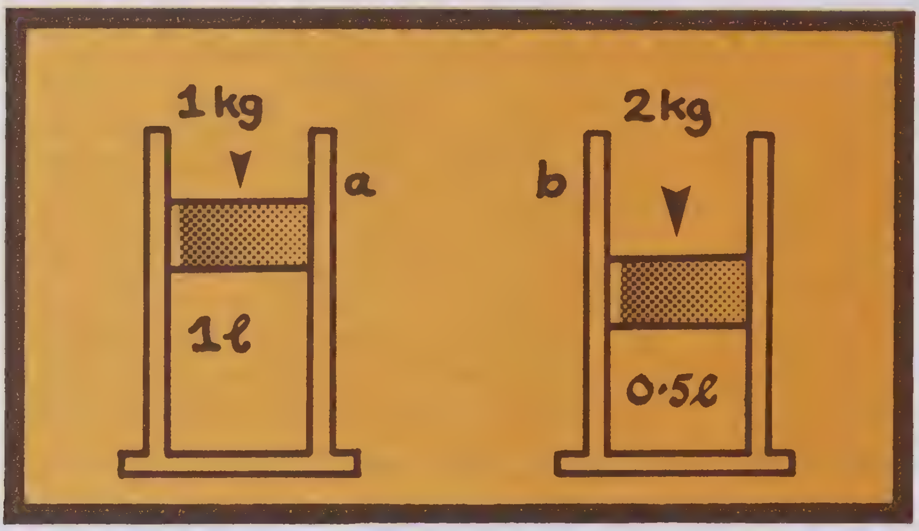
1. The volume of a mass of sulphur dioxide gas at 18°C and 1400 mm pressure is 5.0 ℓ. Calculate its volume at standard conditions (0°C and 760 mm).
2. If a certain mass of gas occupies 1.0 ℓ at STP, what volume will the same mass occupy at 200°C and 300 kPa pressure?
3. Given 1000 ℓ of hydrogen at 15°C and 760 mm. Calculate the volume of this gas at −23°C and 380 mm.
4. Calculate the final volume for each gas given below if the temperature and pressure on the initial volume are changed as indicated.

Initial Volume	*Initial Temperature and Pressure*	*Final Temperature and Pressure*
150 ml	STP	27°C and 152 mm
22.4 ℓ	−23°C and 380 mm	STP
500 ml	20°C and 200 kPa	40°C and 400 kPa
22 400 ml	27°C and 1900 cm	0°C and 76 cm
1.5 ℓ	273°C and 15 kPa	273 K and 200 kPa

5. (a) If 0.60 g of a certain gas with a volume of 400 ml are at 20°C and 500 mm pressure, what will be the volume of this gas at STP?
   (b) What will the mass of the gas be at STP?
   (c) What will the volume be if the temperature and pressure are changed to 40°C and 5000 mm pressure?
6. A certain gas with a volume of 1000 ml at 25°C and 800 mm pressure has a mass of 1.2 g. What will the mass of 500 ml of this gas be if measured at −13°C and 400 mm pressure?

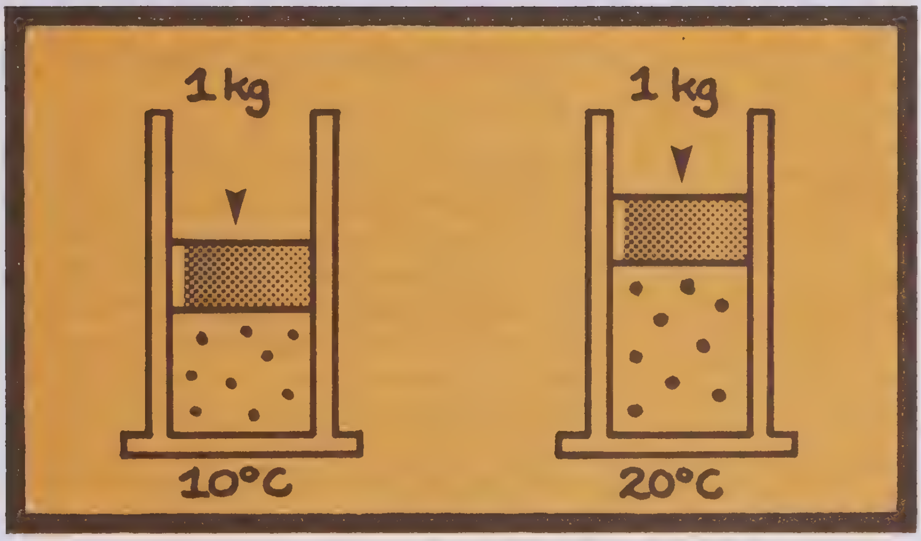
Gas Laws and Particle Model

At the beginning of this unit a particle model was put forward to explain the nature of matter. It is now possible to explain the qualitative aspects of Boyle's and Charles' laws in terms of this model. If the pressure increases the volume decreases.



In the diagram a constant pressure of 1 kg is exerted on the confined gas. The gas colliding with the piston exerts a counterbalancing force equivalent to 1 kg. If the pressure on the top is increased, the pressure of the gas must also increase. This can be done by pushing the particles closer together in the cylinder. There will be more collisions because the travelling distances are shorter. Logically, an increase in pressure on a confined gas would mean a decrease in volume. Boyle's Law can be explained in terms of our particle model.

If the temperature is increased the volume of a given mass of gas increases as long as the pressure remains constant. As the temper-



ature is increased, the pressure exerted by the gas is also increased because the particles are going to move more rapidly as they gain energy from the heat. But the pressure must remain constant. This is done by an increase in volume. As the volume increases the particles must move further before they collide with the piston. The pressure therefore remains constant. Logically, an increase in temperature, if pressure remains constant, means an increase in volume. Thus, Charles' Law can also be explained in terms of our particle model.

In summary you have found that:

1. Equal volumes of different gases have different masses (measured under equal temperatures and pressures).
2. An increase in temperature on a given mass of a confined gas results in an increase in volume as long as the pressure remains constant.
3. An increase in pressure on a given mass of a confined gas results in a decrease in volume as long as the temperature remains constant.

A theory or hypothesis about the nature of gases was proposed and you have tested it. The theory supported the tests. That is, the theory is able to explain the various tests. This is all that any good theory has to do; account for the facts.

Properties of Gases

The results of the previous three experiments can be explained in terms of the particle model of matter. The next two experiments will investigate the nature of the particles. Are the particles of all gases the same or are they different? Do the particles of different gases have similar masses or are they different?

2.7 Properties of Gases

This may be done as a demonstration.

Collect samples of various gases that are available around the lab in 250 ml erlenmeyer flasks and immediately stopper each. Gases that may be used are: oxygen, hydrogen, chlorine, nitrous oxide, nitrogen dioxide, ammonia, hydrogen chloride and carbon dioxide. As you examine these gas samples keep in mind that you are trying to see if the particles of each are the same or different. For example, are any of the gases different in colour? If so, set them aside. Are any of the remaining gases different in their effect on moist litmus? Set them aside. Are any of the remaining gases different in the effect they have on a glowing splint? Set them aside. Can any of the remaining gases be set aside because of some other property that has not already been mentioned? What is that property?

Questions

1. What conclusion can you draw about the particles of different gases, based on the evidence that you gathered in this experiment?

Temperature

Temperature is a measure of how hot an object is. One object is hotter than another if, when the two are brought together, heat flows from one to the other. It is apparent then that hot or cold are relative terms.

Heat is a form of energy; a hot object has more energy than a cold object of the same material and size. If the object referred to is a gas, then the particles of the hot gas have more energy (kinetic energy) than the particles of a cooler gas.

A thermometer is a device for measuring how hot or cold something is. Since temperature is a measure of the average kinetic energy of the particles of a substance, a thermometer is a means of determining this average kinetic energy. For example, with a thermometer the temperatures in two different flasks of gas are found to be 20°C and 40°C. The two numbers tell us that the average kinetic energy of the particles in the second flask is greater than the average kinetic energy of the particles in the first flask.

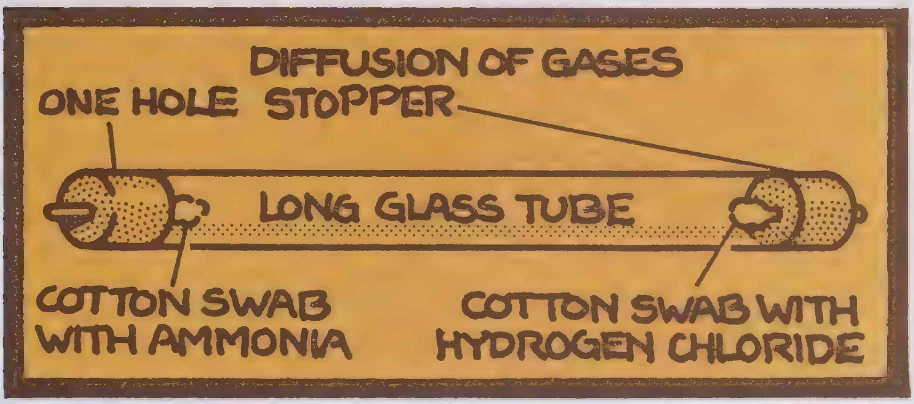
In physics, the kinetic energy of a particle is given by the formula $\frac{1}{2}mV^2$, where m is the mass of the particle and V is its velocity. In the next experiment you will determine the relative velocities of two different gases.

2.8

Diffusion of Gases

retort stand and clamp
120 cm glass tube with 10 mm bore
2 corks
2 droppers
concentrated hydrochloric acid
concentrated aqueous ammonia
2 one-holed rubber stoppers
2 cotton swabs

1. Mount a 120 cm, clean, dry glass tube with a 10 mm bore in a horizontal position on your retort stand with a retort clamp.
2. Measure from the desk to the ends of the tube to make certain that the tube is perfectly horizontal.



3. Break the end off the two cotton swabs close to one end. Slip the stick of the swab through the hole in the stopper. Dip one swab in concentrated hydrochloric acid and the other in ammonia solution. At the same moment insert the two swabs and stoppers into the open ends of the tube.
4. Moisten the stopper of the ammonia bottle with ammonia solution. Bring it near the open mouth of the hydrochloric acid bottle. Hydrochloric acid is a source of hydrogen chloride gas. Ammonia solution is a source of ammonia gas.
5. When the gases meet a white ring forms. Measure the distance from the acid swab to the ring and measure the distance from the ammonia swab to the ring.

Questions

1. Since the particles of ammonia gas and the particles of hydrogen chloride gas were in motion for the same length of time, the distance travelled by each shows their relative velocities.

Which of the two gases travelled more rapidly? Do your classmates agree?

2. Since the temperature of each gas was the same, the average kinetic energy (Ke = $^1/_2 mV^2$) of their particles must be the same. Knowing this, what statement can we make about the relative masses of the particles of each gas?
3. In light of this new information do you have to alter or change your theory about gases and their behaviour?

Gas Model

The last question asked at the end of Experiment 2.4 (Masses of Equal Volumes of Different Gases) was, "How do you account for the fact that the masses of equal volumes of different gases are in fact different?" Turn back in your notebook and see what your response to that question was. Answers that are often proposed are:

1. Equal volumes of different gases at the same temperature and pressure contain the same number of particles, but their masses are different.
2. Equal volumes of different gases at the same temperature and pressure contain different numbers of particles but their masses are the same.

In the analysis of various samples of different gases some of the information gathered would indicate that the particles of different gases are in fact different. In that experiment it was shown that the particles of various gases may differ in colour, odour, effect on litmus, effect on a glowing splint or reaction with air. In the diffusion experiment, however, something much more significant was illustrated. The two gases diffused at different rates even though the temperature of each was the same. Just what does this mean?

Since the temperature of each gas was the same, the average kinetic energy of the particles must have been the same.

Gas 1 (Hydrogen Chloride)	Gas 2 (Ammonia)
Ke = $^1/_2 m_1 V_1^2$	Ke = $^1/_2 m_2 V_2^2$

But the two kinetic energies are the same.

$$\therefore\ ^1/_2\, m_1 V_1^2 = {}^1/_2\, m_2 V_2^2$$

$$m_1 V_1^2 = m_2 V_2^2$$

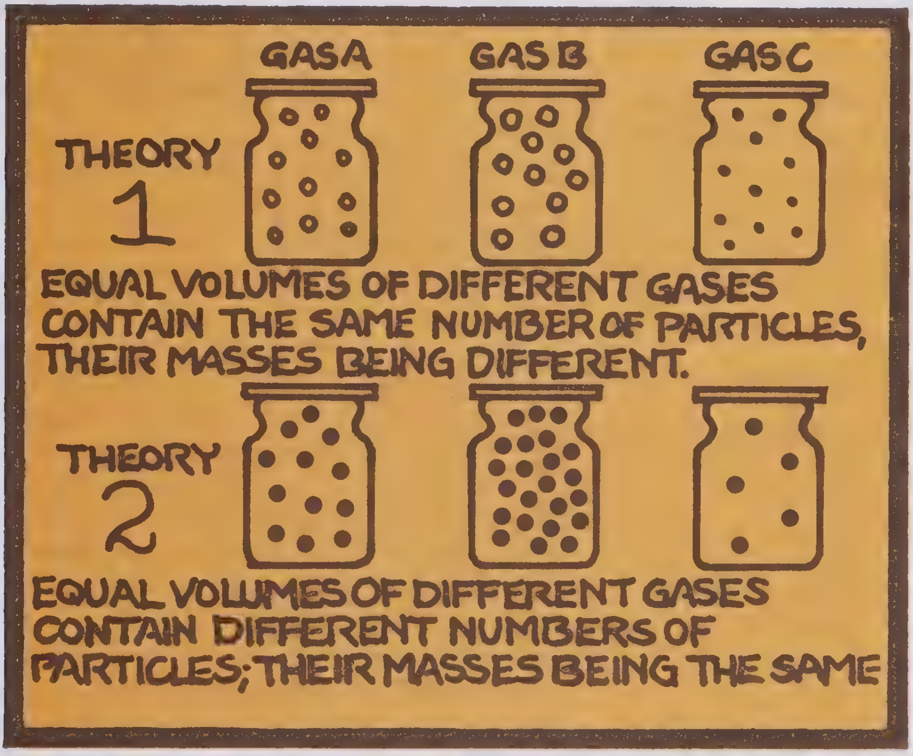
$$\text{or } \frac{V_1^2}{V_2^2} = \frac{m_2}{m_1}$$

But V_2 was greater than V_1, therefore m_1 must be greater than m_2.

The masses of the particles of the two gases are different.

If the two gases selected are representative of gases in general then it would be safe to say that the particles of different gases have different masses.

Now, the question that arises is how does this help us with the problem of masses of equal volumes. If the masses of the particles are different then it would seem that equal volumes of different gases at the same temperature and pressure contain equal numbers of particles but their masses are different. It might seem unbelievable, but all the evidence supports such a proposal. The next reasonable step to take is to test such a proposal. Can it explain the results of other experiments?

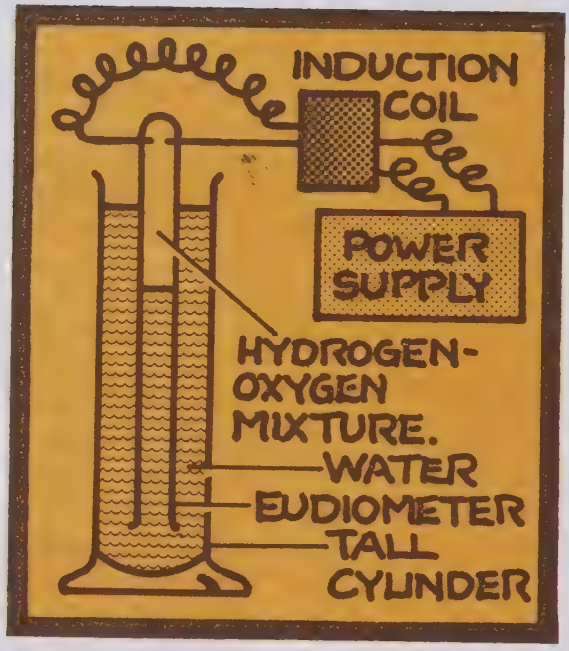

GAS A
GAS B
GAS C
THEORY 1
EQUAL VOLUMES OF DIFFERENT GASES CONTAIN THE SAME NUMBER OF PARTICLES, THEIR MASSES BEING DIFFERENT.
THEORY 2
EQUAL VOLUMES OF DIFFERENT GASES CONTAIN DIFFERENT NUMBERS OF PARTICLES; THEIR MASSES BEING THE SAME

2.9 Combustion of Hydrogen (Demonstration)

Warning
This experiment is extremely dangerous and should only be attempted as a demonstration and then only when adequate safety precautions are taken.

eudiometer
tall cylinder
induction coil
6 V power supply
cedar splints
gas bottles

source of hydrogen gas
source of oxygen gas



1. Fill a eudiometer with water and immerse it in a tall cylinder filled with water (see diagram).
2. Bubble about 10 ml of hydrogen gas into the eudiometer. Measure the volume of the gas accurately while the level of the water is the same inside and outside the eudiometer.
3. Bubble an equal volume of oxygen gas into the eudiometer. Read the total volume of gas in the tube but once again take the reading with the water level the same inside and outside the eudiometer. Allow the gas mixture to stand for about 10 min. Why?
4. Lower the eudiometer into the tall cylinder of water until only the electrodes are above the surface. Wrap the bottom of the tube in a towel exposing only the electrodes. Spark the gas mixture with an induction coil or a Tesla coil.
5. Allow the eudiometer to cool, then raise the tube in the tall cylinder until the water level is the same in the eudiometer as in the cylinder. Read the final volume of gas in the tube.
6. Raise the eudiometer from the cylinder of water with the open end blocked with your finger to prevent the water from escaping. Turn the eudiometer upright and bring a glowing splint to the gas remaining in the tube.

Table 2.8

Volume of hydrogen used ______ ml
Volume of the gas mixture ______ ml
Volume of oxygen used ______ ml
Volume of gas remaining ______ ml

7. Collect a gas bottle full of each of the gases used in the experiment. Bring a glowing splint to the mouth of each.
8. Record your observations in a table like Table 2.8

Questions

1. Why is it necessary to use a tall cylinder rather than a beaker to hold the eudiometer? Give at least two reasons.
2. What gas remained in the eudiometer tube after the reaction had taken place?
3. Copy the following statements into your notebook and complete them:
   (a) ____________ ml of hydrogen reacted with ____________ ml of oxygen.
   (b) ____________ volumes of hydrogen react with ____________ volume of oxygen.

Before attempting any explanation for the volume of hydrogen combining with oxygen, let us try just one more experiment to see if other gases combine in some simple ratio like hydrogen and oxygen did.

This time we will choose gases that, with some care, may be handled by everyone and the experiment may be done as a class experiment.

2.10

The Reaction Between Nitric Oxide and Oxygen

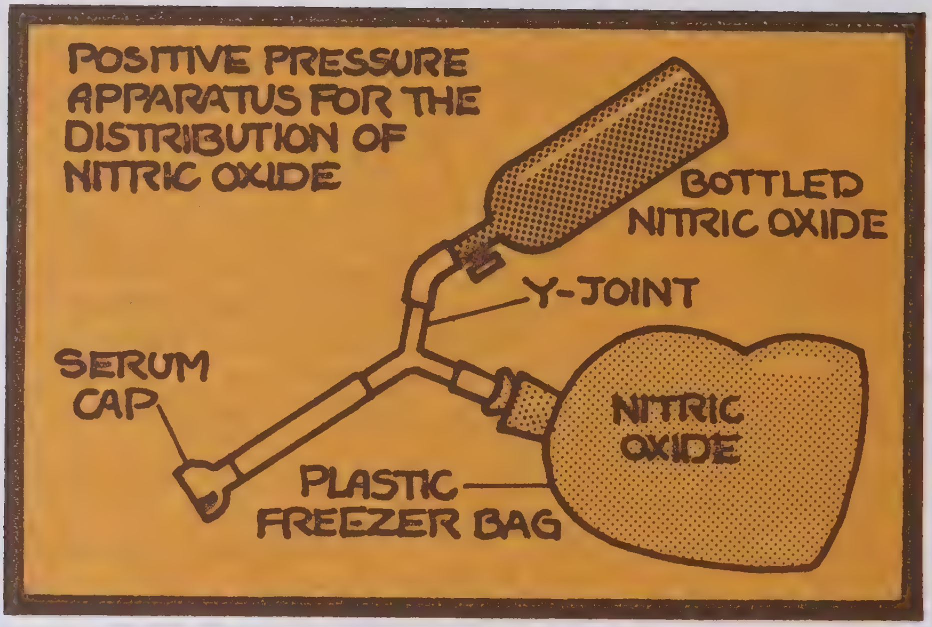
Warning

Do not inhale the gases used in this experiment. This experiment should only be done in a well-ventilated room.

60 cm length of 11 mm glass tubing
funnel
syringe
serum cap
large graduated cylinder
rubber stopper

source of oxygen gas
source of nitric oxide gas

1. Stopper a 60 cm length of 11 mm glass tubing with a serum cap. Using a funnel, fill the tube with water and invert the open end


POSITIVE PRESSURE APPARATUS FOR THE DISTRIBUTION OF NITRIC OXIDE
BOTTLED NITRIC OXIDE
Y-JOINT
SERUM CAP
NITRIC OXIDE
PLASTIC FREEZER BAG

into a large graduated cylinder of water. Be sure to keep the open end of the tubing below the surface of the water.

2. Add 20 ml of oxygen to the tube using your syringe. Mark the level of the water in the tube with a piece of tape or a marking pen.
3. Draw 10 ml of NO gas into the syringe from the source supplied by your teacher. Seal it by sticking the needle into a rubber stopper. Inject it through the cap into the oxygen. Record all observations carefully.
4. Being very careful to keep the open end below the surface of the water, agitate the tube gently. When you are sure the water has stopped rising, inject air into the tube with a syringe. Record the volume of air required to return the water to the mark in Step 2.
5. Record your observations in a table like Table 2.9

Table 2.9

Volume of oxygen added	__________ ml
Volume of nitric oxide added	__________ ml
Volume of air required in Step 4 (Volume of oxygen used up)	__________ ml

Calculate the ratio:

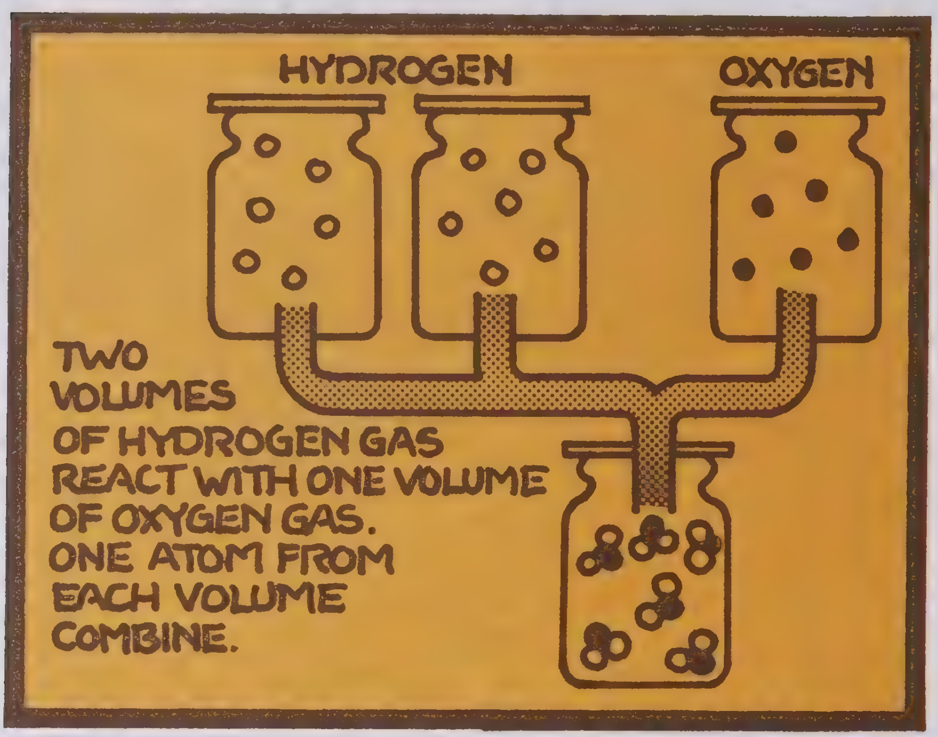
$$\frac{\text{Volume of nitric oxide added}}{\text{Volume of oxygen used}}$$

6. Repeat Steps 3, 4 and 5 twice more and average your ratios. Complete the following statements in your notebook using averaged results:
   (a) 10 ml of nitric oxide react with __________ ml of oxygen.
   (b) __________ volumes of nitric oxide react with __________ volumes of oxygen.

Up to this point, we have gathered evidence that supports the particle theory of gases. Further evidence indicates that the particles of various gases differ in physical and chemical properties as well as in mass. But the most amazing piece of evidence gathered indicated that equal volumes of two different gases at the same temperature and pressure contain equal numbers of molecules. As difficult as this is to digest, assume for a moment that the information is correct and see if it agrees in any way with the results of the last two experiments.

The simplest and most basic explanation that we can offer for these reactions is that one particle from each volume of one gas reacts with one particle from each volume of the other gas used.

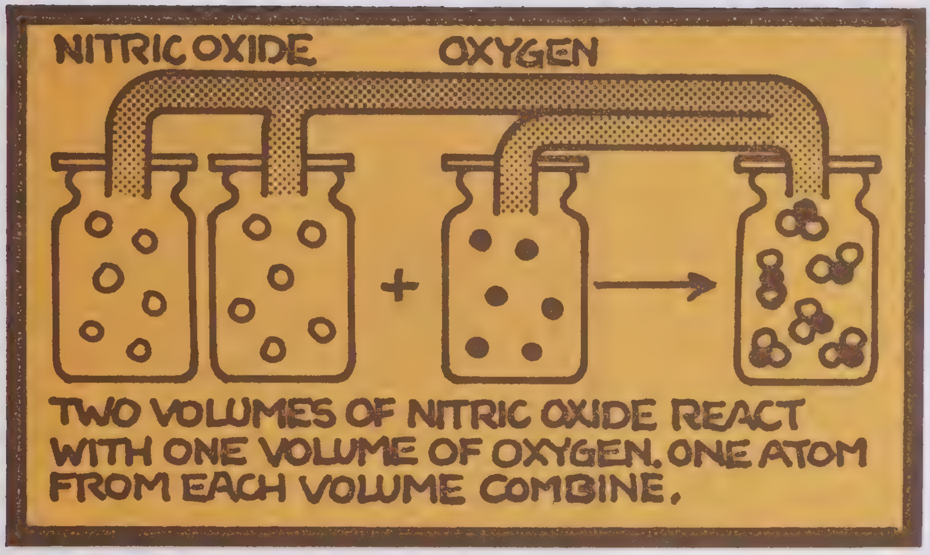
Hydrogen reacts with oxygen:
2 volumes ——→ 1 volume


HYDROGEN
OXYGEN
TWO VOLUMES OF HYDROGEN GAS REACT WITH ONE VOLUME OF OXYGEN GAS. ONE ATOM FROM EACH VOLUME COMBINE.

Since two volumes of hydrogen reacted with one volume of oxygen then two particles of hydrogen (one from each volume) reacted with one particle of oxygen. If equal volumes of the gases contained equal numbers of particles, this would mean that when all the hydrogen from the two volumes was used, all the oxygen from one volume would also be used. The results of this experiment can be explained on the basis of equal volumes of different gases containing the same numbers of particles.

Nitric oxide reacts with oxygen:
2 volumes ——→ 1 volume

Similarily the results of this experiment can be explained on the basis of equal volumes containing the same number of particles and the one particle from each volume ratio.


NITRIC OXIDE
OXYGEN
+
TWO VOLUMES OF NITRIC OXIDE REACT WITH ONE VOLUME OF OXYGEN, ONE ATOM FROM EACH VOLUME COMBINE.

A theory must be able to explain the facts. The theory that equal volumes of different gases (measured at the same temperature and pressure) contain equal numbers of particles does an excellent job of explaining the facts that were gathered in the last two experiments. This is the same theory that was proposed in the nineteenth century by Amedeo Avogadro and to this day the theory still holds true. The theory just stated is generally referred to as Avogadro's Hypothesis.

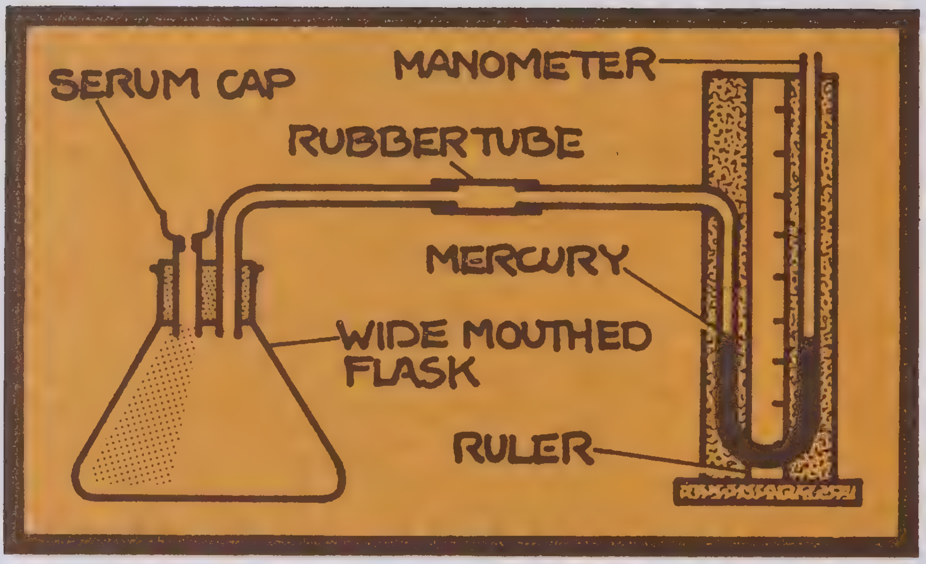
Pressure in a Mixture of Gases

It appears from the evidence presented in the last series of experiments that gases are made up of particles and that these particles are different in mass and nature for each different gas. Further, the smaller the particle, the greater will be its velocity at any given temperature. Gas pressure is the result of collisions between the moving particles and the container they are in. The magnitude of the force exerted by any single collision will depend on two things. The first is the mass of the particle, which is like catching a ping-pong ball or a baseball hurled at you with the same velocity. The second thing is the velocity of the particle, which is like catching a baseball hurled overhand or one gently thrown underhand.

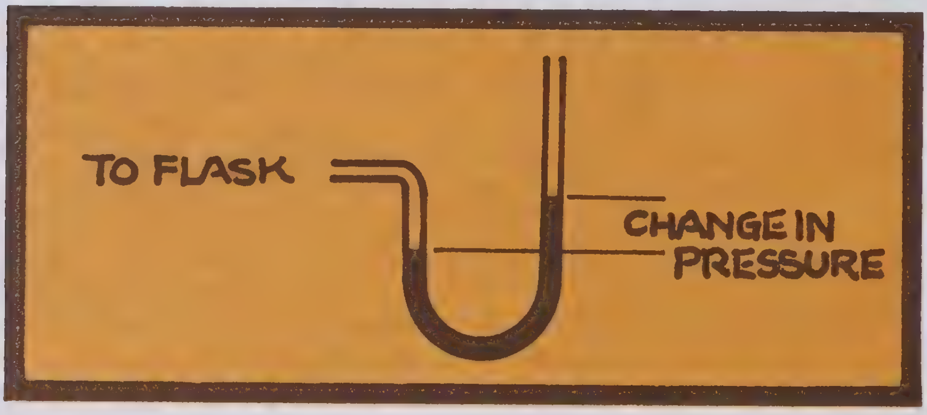
2.11

Partial Pressure

side-arm flask
manometer
serum cap assembly (see diagram)
ruler
10 ml syringe and needle
various gases



1. Set up the apparatus as shown in the diagram. All connections must be airtight and the mercury level must be the same in each arm of the manometer. This can best be done by inserting the needle through the serum cap and adding or removing a little bit of air as required.
2. Measure the height of the mercury in each arm of the manometer. This measurement may be made from the desk (see diagram).



3. Put 10 ml of one of the gases provided by your teacher in the syringe. Add this amount of gas slowly to the flask then immediately remove the syringe from the serum cap. Measure and record the height of the mercury in each arm of the manometer now.
4. Loosen the stopper to allow the mercury to return to its original position. Tighten the stopper and again make certain that the mercury level is the same in each arm of the manometer (see Step 1).
5. Repeat Steps 3 and 4 with each of the other gases available. All measurements must be made with maximum care.

Questions

1. After the mercury levels were made the same in each arm of the manometer what was the pressure on the mercury in the open end of the manometer?
2. What must the pressure have been in the flask at this point? Explain.
3. Why did the mercury level change when a gas was added? Explain.
4. Explain the magnitude of the pressure change found for each of the various gases used.
5. Are the results of this experiment consistent with:
   (a) the particle theory of gases?
   (b) Avogadro's Hypothesis? Explain.

Molecules and Atoms

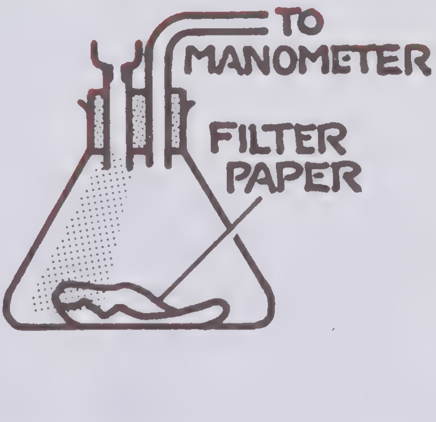
The particle theory of matter has been so reliable over the years that chemists have given these particles a name; they are called *molecules*. Molecules are tiny fundamental particles of which many substances are made. It would be reasonable at this point to go back through the unit and replace the word "particle" with the word "molecule."

In the sublimation experiment, solid iodine turned to a purple vapour and then returned to a solid, and solid benzoic acid seemed to disappear then reappeared through a filter paper. In each case the solid changed to molecules then back to a solid again. It would seem reasonable to assume that the solids themselves were made up of the molecules but they were just more closely packed.

2.12

Pressure of Liquids

manometer
syringe and needle
250 ml flask
liquids such as water,
methanol and acetone



1. Set up the apparatus exactly as in Experiment 2.11, with this one difference. Before placing the stopper in the flask add a piece of filter paper or paper towel (see diagram).
2. Equalize the pressure on each side of the manometer by either adding or removing a small amount of air with a dry syringe.
3. Place several millilitres of one of the liquids to be used in your syringe. Add a few drops of the liquid through the serum cap to the paper in the flask. Add enough liquid so that the paper in the flask appears slightly damp. Remove the syringe.
4. After several minutes, measure the difference in height of the mercury on each side of the manometer. Record this number in your record book.
5. Remove the stopper from the flask, remove the paper and dry the flask. Add a new piece of paper and repeat Steps 1 to 4 with a second and a third liquid as directed by your teacher.

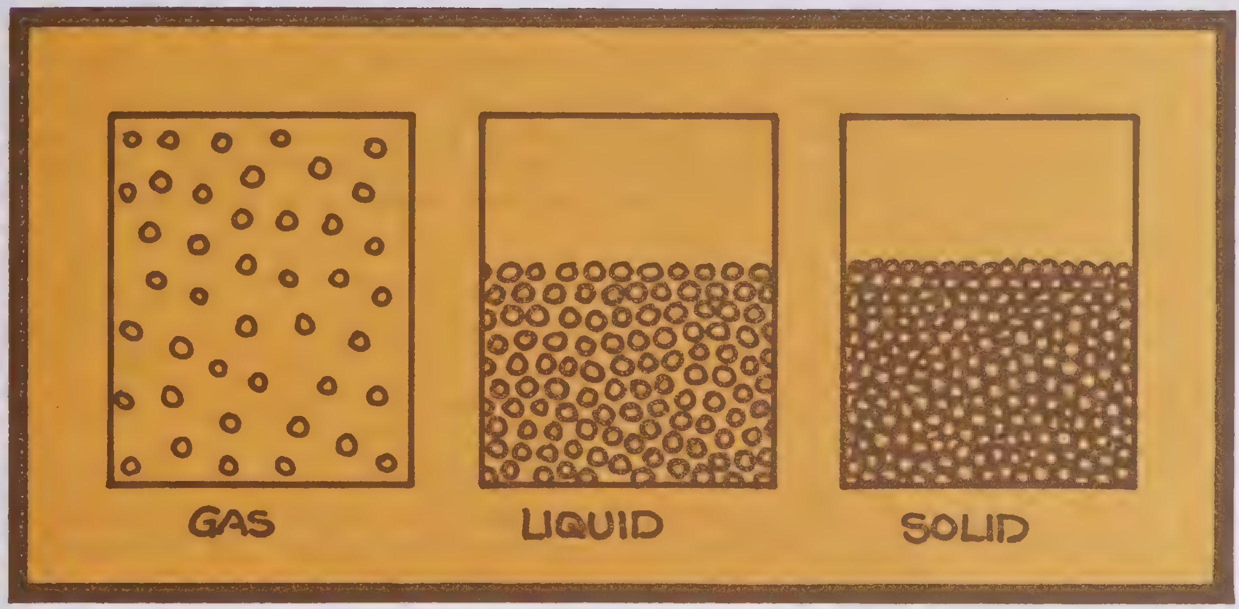
Questions

1. Suggest a reason for putting the paper towel in the flask.
2. Does the change in the level of the liquid in the manometer arms indicate an increase or a decrease in pressure in the flask?
3. Why must enough liquid be added to moisten the paper?
4. If temperature is kept constant in the flask how do you account for the increase in pressure?
5. Did all liquids used cause the same change in pressure? How could any differences be explained?

Vapour Pressure

Could it be that liquids turn into gases under certain conditions? Perhaps when a liquid was added to the flask, some of it became a gas, thus increasing the number of molecules in the flask. An increase in the number of molecules would mean a greater number of collisions within the flask and therefore an increase in the pressure. This would be shown by a change in the liquid level in the manometer arms. The pressure exerted by a vapour from a liquid is called the *vapour pressure* of the liquid.

What is most significant about this experiment is that the results can be explained by assuming that part of the liquid becomes a vapour. The pressure increases up to a point then appears to stop. Does this mean that only a portion of the liquid becomes gaseous or is capable of becoming gaseous? It is more probable that as the gaseous molecules move about in the flask, some of them turn into a liquid again. A point is eventually reached where the number of



molecules leaving the liquid to become a gas is equal to the number leaving the gaseous state to become a liquid. At this point the pressure in the flask appears to remain constant.

If the liquids used are capable of becoming gaseous molecules then it may not be unreasonable to assume that liquids themselves are made up of molecules. Since liquids are free to pour and take the shape of the container they are in, it would seem that the molecules of a liquid are more loosely packed than those of solids and have greater freedom of movement.

Earlier in this course you discovered that certain substances could be decomposed or broken down into simpler substances. For example, when cobalt chloride, copper sulphate or cupric carbonate were heated two new substances were formed from the

John Dalton

molecules of the original substance. How is this possible if the molecule is the smallest, most fundamental unit of matter? Chemists have an answer to this question. They have proposed that all matter, regardless of form, is built up of a tiny basic unit called an *atom*. No atom can be formed from simpler substances nor can an atom be broken down into simpler substances.

The Elements

There are about 104 different kinds of atoms called elements, each with its own name. An element is a substance that cannot be broken into simpler substances and forms the basis of all matter. Every known substance on this earth is made up of the elements or some combination of them.

In 1814, Jöns Berzelius suggested that the elements should be given a symbol based on their names and the now-accepted system of using a one- or two-letter symbol was introduced.

Element	*Symbol*	*Element*	*Symbol*
Oxygen	O	Iodine	I
Hydrogen	H	Yttrium	Y
Carbon	C	Vanadium	V
Fluorine	F	Nitrogen	N
Boron	B	Phosphorus	P
Sulphur	S	Uranium	U

The symbol for other elements is a capital letter followed by a small letter.

Element	*Symbol*	*Element*	*Symbol*
Helium	He	Silicon	Si
Lithium	Li	Chlorine	Cl
Beryllium	Be	Argon	Ar
Neon	Ne	Calcium	Ca
Aluminum	Al	Manganese	Mn

From your point of view there is one complicating factor and this is that not all symbols are derived from the English name. In some cases the symbol is derived from the Latin name.

Element	*Symbol*	*Origin*
Iron	Fe	Ferrum – Latin
Lead	Pb	Plumbum – Latin
Gold	Au	Aurum – Latin
Silver	Ag	Argentum – Latin
Antimony	Sb	Stibium – Latin
Tin	Sn	Stannum – Latin
Sodium	Na	Natrium – Latin
Potassium	K	Kallium – Latin
Copper	Cu	Cuprum – Latin
Tungsten	W	Wolfram – German

Refer to the table or chart at the end of this book for a complete list of the elements and their symbols.

The Mole

Chemistry is more that just a study of how things behave under certain conditions, even though it is always interesting to see what happens to something when it is heated or what will happen if you mix two different things together. Chemists have found it much more useful to put quantities on things and find out how much of a new substance may be produced. In the synthesis of hydrogen and oxygen we found something out about how important quantities are. In that experiment, two volumes of hydrogen reacted with one volume of oxygen. It would be more useful to know the number of particles (molecules) of hydrogen and the number of particles of oxygen that reacted with each other. But since the particles are atoms or molecules and the period at the end of this sentence may contain billions of such particles, the numbers would be extremely large. To avoid having to use very large numbers chemists have introduced a new unit to chemistry called the *mole* (mol). A mole is simply a large number of things; the number in fact is 6.02×10^{23}.

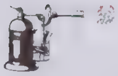
$$1 \text{ mol} = 6.02 \times 10^{23} \text{ things.}$$

6.02×10^{23} atoms is 1 mol of atoms;
6.02×10^{23} pencils is 1 mol of pencils;
6.02×10^{23} dollars is 1 mol of dollars.

The value of such a unit is shown in Experiment 2.13.

2.13

Glass Bead Experiment



250 ml beaker	glass beads
balance	graduated cylinder

1. Read through the experiment and draw up a data sheet in your notebook.
2. Find the mass of a clean, dry 250 ml beaker.
3. Count 50 glass beads and place them in the beaker. Find the mass of the beaker and the beads.
4. Calculate the mass of 50 beads.
5. Calculate the mass of 1000 beads.
6. Determine the mass of 1000 beads and the 250 ml beaker used above.
7. Adjust the balance to the combined mass of beaker and 1000 beads. Place the beaker on the balance pan and add beads until the balance is balanced with the beaker and beads.
8. How many beads should now be in the beaker?
9. Measure the volume of the beads by pouring them into a graduated cylinder.

Questions

1. What is the volume of 1000 beads?
2. What is the mass of 10 000 beads?
3. What is the volume of 10 000 beads?
4. Suggest two ways of "counting out" large numbers of glass beads.
5. A mole of molecular oxygen gas has a mass of 32 g which at 100 kPa pressure and 25°C occupies 24.5 ℓ. Suggest two ways of "counting out" 3.01×10^{23} molecules of oxygen.

Consider for a moment the information that the eudiometer gives us in conjunction with Avogadro's Hypothesis.

$$\begin{array}{lcll} \text{Hydrogen} & + & \text{Oxygen} & \longrightarrow \text{Water} \\ \text{2 volumes} & + & \text{1 volume} & \longrightarrow \text{2 volumes} \end{array}$$

Suppose for a moment that one volume of oxygen contained 1000 molecules. That means that two volumes of hydrogen contain 2000 molecules and two volumes of water contain 2000 molecules.

$$\begin{array}{lcll} \text{Hydrogen} & + & \text{Oxygen} & \longrightarrow \text{Water} \\ \text{2 volumes} & + & \text{1 volume} & \longrightarrow \text{2 volumes} \\ \text{2000 molecules} & + & \text{1000 molecules} & \longrightarrow \text{2000 molecules} \end{array}$$

Each water molecule must have at least one oxygen atom in it so 2000 water molecules must contain at least 2000 atoms of oxygen which came from 1000 molecules of oxygen. That means that each oxygen molecule must contain at least two atoms of oxygen. Oxygen forms a *diatomic molecule*: O_2.

From similar types of reactions involving gases we can find similar evidence for the fact that many gases are diatomic. Among them are:

N_2	Nitrogen	Cl_2	Chlorine
O_2	Oxygen	Br_2	Bromine
F_2	Fluorine	I_2	Iodine

In fact some elements are *triatomic*. Oxygen may be triatomic and is called ozone when it is – O_3.

Ozone is responsible for the sharp clean smell around electrical equipment.

There are some six gases that are *monatomic*. In these gases a molecule consists of a single atom. They are:

He	Helium	Kr	Krypton
Ne	Neon	Xe	Xenon
Ar	Argon	Ra	Radon

Atomic and Molecular Masses

Because of the lack of sophisticated equipment, chemists did not know the actual masses of individual atoms for many years. But this was not really a problem because they did know the relative masses of the atoms. What this means is that they knew how much more or how much less mass each atom had compared to some standard. Just as you could arrange all the students in your class according to mass if you knew how much lighter or heavier they were than yourself, chemists could do the same with the atoms of the various elements. Chemists chose carbon as the standard and assigned it a mass of 12. All other atoms were compared to this standard. No unit was assigned to the numbers because they were only *relative* masses. For example:

Hydrogen — atomic mass = 1.

This means that a hydrogen atom is $\frac{1}{12}$ the mass of the carbon 12 atom.

But with the development of the necessary sophisticated equipment this carbon 12 atom was found to have an actual mass of 1.99×10^{-23} g. This means that a very large number of carbon atoms would be required to amount to a mass of 12 g.

$$1 \text{ carbon atom has a mass of } 1.99 \times 10^{-23} \text{ g}$$

$$x \text{ carbon atoms have a mass of } 12 \text{ g}$$

$$x \; \frac{12 \text{ g}}{1.99 \times 10^{-23} \text{ g}} = 6.02 \times 10^{23} \text{ atoms}$$

You will of course recognise this number as being a mole. Therefore, 1 mol of carbon atoms has a mass of 12 g. The mass in grams of 1 mol of carbon atoms is called its *gram atomic mass*. Similarly, the mass in grams of 1 mol of atoms of any element is the gram atomic mass.

Example 1

1 mol of hydrogen atoms has a mass of 1 g

∴ gram atomic mass of hydrogen = 1 g

Example 2

1 mol of oxygen atoms has a mass of 16 g

∴ gram atomic mass of oxygen = 16 g

Gram Molecular Masses

Earlier in this unit you were given evidence to indicate that oxygen molecules contain at least two atoms (diatomic). This would mean that 1 mol of oxygen molecules would contain 2 mol of oxygen atoms (two gram atomic masses).

1 mol of oxygen atoms (1 GAM) = 16 g
∴ 2 mol of oxygen atoms (2 GAM) = 32 g
∴ 1 mol of oxygen molecules = 32 g

This mass is referred to as the *gram molecular mass* of oxygen gas.

Example 1

Sugar ($C_6H_{12}O_6$) contains 6 carbon atoms
12 hydrogen atoms
6 oxygen atoms per molecule

∴ gram molecular mass = 6 (12 g) GAM of C = 72 g
12 (1 g) GAM of H = 12 g
6 (16 g) GAM of O = 96 g

∴ gram molecular mass = 180 g

Example 2

Water (H_2O) contains two hydrogen atoms and one oxygen atom per molecule.

1 mol of water molecules contain 2 mol of H atoms
= 2 GAM = 2 g
1 mol of water molecules contains 1 mol of O atoms
= 1 GAM = 16 g
∴ gram molecular mass of water = 16 + 2
= 18 g

Problems

Refer to the back of the text for atomic masses.

1. How many grams are there in:
   (a) 1 mol of Na;
   (b) 0.5 mol of Ca;
   (c) 1.25 mol of Al;
   (d) 0.015 mol of Br?

2. How many moles are there in:
   (a) 46 g of Na;
   (b) 152.5 g of Sb;
   (c) 49.25 g of Au;
   (d) 325 g of Zn;
   (e) 2 kg of Hg?
3. Find the gram molecular mass (GMM) of:
   (a) HCl;
   (b) H_2SO_4;
   (c) H_3PO_4;
   (d) $C_{12}H_{22}O_{11}$;
   (e) CH_3COOH;
   (f) C_2H_5OH.
4. How many grams are there in each of the following:
   (a) 1 mol of HCl;
   (b) 0.25 mol of H_2SO_4;
   (c) 2.5 mol of H_3PO_4;
   (d) 0.15 mol of C_2H_5OH?
5. How many moles are there in each of the following:
   (a) 67 g of $CaSO_4$;
   (b) 82 g of Na_3PO_4;
   (c) 175 g of H_2SO_4;
   (d) 25 g of $AlPO_4$;
   (e) 5 kg of $FeSO_4$?

Molar Volume

2.14 The Volume of 1 mol of a Gas at STP

Use the results of Experiment 2.13.

1. Collect the following information from Experiment 2.13 and record it in a table like Table 2.10.

Table 2.10

Gas Used	*Pressure (mm)*	*Temp. (°C)*	*Actual Mass*	*Volume of Bag*
(i)				
(ii)				
(iii)				

Calculations

1. For each gas used, determine what the actual volume would have been if measurements had been made at STP. (Use Boyle's and Charles' Laws).
2. Use the formula of the gas used in each case to determine the gram molecular mass.
3. Use GMM and actual mass of each gas to determine how many moles of each gas were used in each case.
4. Use moles of gas used and volume at STP to determine the volume of 1 mol of each gas measured at STP.
5. Since 1 mol of different gases represents the same number of molecules, according to Avogadro's Hypothesis the volumes for 1 mol of each gas should be the same. How do the volumes that you have calculated compare? Compare class results. Use a reference book to determine the accepted molar volume measured at STP.
6. The volume that you have just determined is called the *molar volume*, or sometimes the gram molecular volume. Write out a suitable definition for molar volume.
7. The calculations that you have just completed and Experiment 2.13 suggest a means of determining experimentally the gram molecular mass for a gaseous substance. Describe an experiment and the necessary calculations that would have to be done to determine a gram molecular mass for a gaseous substance.

Ideal Gas Law

Boyle's Law may be expressed mathematically as:

$$V \alpha \frac{T}{p} \text{ or } V = \frac{k}{p} \text{ (temperature is constant)}$$

Charles' Law may be expressed mathematically as:

$$V \alpha T \text{ or } V = kT \text{ (pressure is constant)}$$

Since the two laws are concerned with the same three things, volume, pressure and temperature, they can be combined and expressed as

$$V = k\frac{T}{p}$$

for a given mass of gas. The ideal gas is one that would obey these laws and conform exactly to Avogadro's Hypothesis. In fact, however, no gas will do this. Many gases come very near to satisfying these conditions over small variations in temperature and pressure that are near to standard conditions. Also, to conform to Avogadro's Hypothesis the volume (at constant temperature and pressure) must be dependent on the number of molecules.

$$\text{Therefore, } V \; \alpha \; \text{n} \times \frac{T}{p}$$

This means that we must introduce a new proportionality constant. R is the symbol for the constant most commonly used.

$$\text{Hence, } V = \text{Rn}\,\frac{T}{p}$$

This equation can be rearranged to give:

$$pV = \text{nR}T$$

The value for R depends on the quantities selected for 1 mol of a gas.

$$\text{Thus, R} = \frac{pV}{\text{n}T}$$

$$= \frac{100 \text{ kPa} \times 22.4\,\ell}{1 \text{ mol} \times 273 \text{ K}}$$

$$= 8.2\;\frac{\ell - \text{kPa}}{\text{K} - \text{mol}}$$

Using the Ideal Gas Law

Example 1

What volume would 0.5 mol of oxygen gas measured at 27°C and 800 mm pressure occupy?

In finding the value for R the pressure was expressed in kilopascals.

$$\therefore 800 \text{ mm} = \frac{800 \times 100}{760} = 105 \text{ kPa}$$

$$\text{and } 27°\text{C} = 273 + 27 = 300 \text{ K}$$

$$\therefore V = \frac{\text{nR}T}{p} = \frac{0.5 \text{ mol} \times 300 \text{ K} \times 8.2}{105 \text{ kPa}}\;\frac{\ell - \text{kPa}}{\text{K} - \text{mol}}$$

$$= 11.71\,\ell = 1 \text{ x } 10\ell$$

Example 2

100 mol of chlorine gas at STP are to be stored at 2500 kPa pressure and 20°C. What will be the volume of the chlorine?

$$V = \frac{\text{nR}T}{p} = \frac{100 \text{ mol} \times 293 \text{ K} \times 8.2}{2500 \text{ kPa}}\;\frac{\ell - \text{kPa}}{\text{K} - \text{mol}}$$

$$= 96.1\,\ell$$

Summary

Chemical change can be recognized by a change in colour, heat absorbed or released, a gas released, or the formation of a precipitate. Further, chemical changes are characterized by a constancy of mass since they obey the Law of Conservation of Mass. Also, when compounds are formed by chemical reactions, they conform to the Law of Definite Proportions.

Matter, whether it is solid, liquid or gas, is made up of particles that are either atoms or molecules. All matter as we know it is made up from the 104 known elements either by themselves or combined in various proportions as compounds.

Gas volumes change with pressure and temperature changes, as shown by Boyle's and Charles' Laws. Further, each gas in a mixture of gases appears to react independently of any other gas present. This was illustrated by the fact that the partial pressure (pressure of a gas in a mixture of gases) of each different gas was exactly the same for each 10 ml sample selected and used under the same conditions. This means that if 100 ml of each of three different gases were placed in a closed system and the total pressure was 300 kPa, then the partial pressure of each gas would be 100 kPa. Further, if 100 ml of a fourth gas that would not react with any of the first three gases were added, the total pressure would rise to 400 kPa. This idea of the additivity of the partial pressures of various gases was first proposed by John Dalton and is referred to as Dalton's Law of Partial Pressures.

Equal volumes of gases measured at the same temperature and pressure contain equal numbers of molecules. If the volumes referred to are 22.4 ℓ measured at STP then 1 mol (6.02×10^{23}) of molecules will be in each volume. The mole is a unit widely used by chemists because of the vast numbers of particles in the smallest of quantities. The volume, 22.4 ℓ, measured at STP is called the molar volume.

All matter, whether a solid, a liquid or a gas, consists of particles called atoms which in turn form molecules. In solids and liquids the particles are close together, while in gases they are so far apart they act independently of each other. This is known as the atomic theory. No other theory in chemistry is or has been as important. It is a theory that has withstood the test of time and will be used often throughout this course.

A theory, however, is only as good as the tests it is put to. In the past few experiments, we have tested our particle theory of matter on numerous occasions. Each test was more of a challenge to the theory than the previous one. Each time, however, the theory withstood the test.

Perhaps we have only selected simple tests and avoided those that our theory would fail. What test might you use to test our theory? Can you think of any change in matter that cannot be explained in terms of particles?

Recall how we defined a solid, a liquid and a gas. Also, remember that any given substance may exist in each of the three states. That is, the same particles must make up the solid state of a substance as well as its liquid and its gaseous states. The difference must, therefore, be in the arrangement of the particles. Can you explain the difference between the three states of a particular substance in terms of particles? Perhaps a drawing of your mental picture of each would be better than a written description. Can you think of a better theory to explain the structure of matter and its changes?

Questions and Problems

1. (a) A roll of steel wool with a mass of 1.05 g is ignited, cooled and its mass determined again. After burning, the residue had a mass of 1.5 g. Does this information contradict the Law of Conservation of Mass? Explain.
   (b) In a second trial 1.52 g of steel wool were ignited. How much residue would be produced?
   (c) How many moles of iron (Fe) were used in each case?
   (d) If the residue is Fe_2O_3, how many moles of it were produced in each case?
2. When a sample of cupric oxide with a mass of 5.925 g was heated over a stream of hydrogen gas, 4.74 g of copper metal were produced. When 5.45 g of copper metal were made to react completely with 6.81 g of oxygen, 12.26 g of cupric oxide were produced. Show how the results of these experiments illustrate the Law of Definite Proportions.
3. If hydrochloric acid is poured over marble chips a clear colourless gas is produced along with a yellowish coloured solution. If the solution is evaporated to dryness a whitish solid remains. Does this sound like a decomposition reaction or a synthesis reaction? Explain.
4. How does the particle model of matter explain the Law of Definite Proportions?
5. Why does a hydrogen-filled balloon float in air?
6. Why are gases compressible?
7. A gas bag as used in Experiment 2.5 is filled with 0.62 ℓ of hydrogen gas (H_2) measured at 27°C and 750 mm pressure. The actual mass of the gas in the bag was found to be 0.05 g. Use this information to calculate the molar volume of hydrogen.
8. How many molecules would there be in a 5 ℓ sample of a gas that was measured at -23°C and 200 kPa?
9. How many atoms are there in 1 mol of each of the following:
   (a) HCl;
   (b) H_3PO_4;
   (c) CH_3COOH?

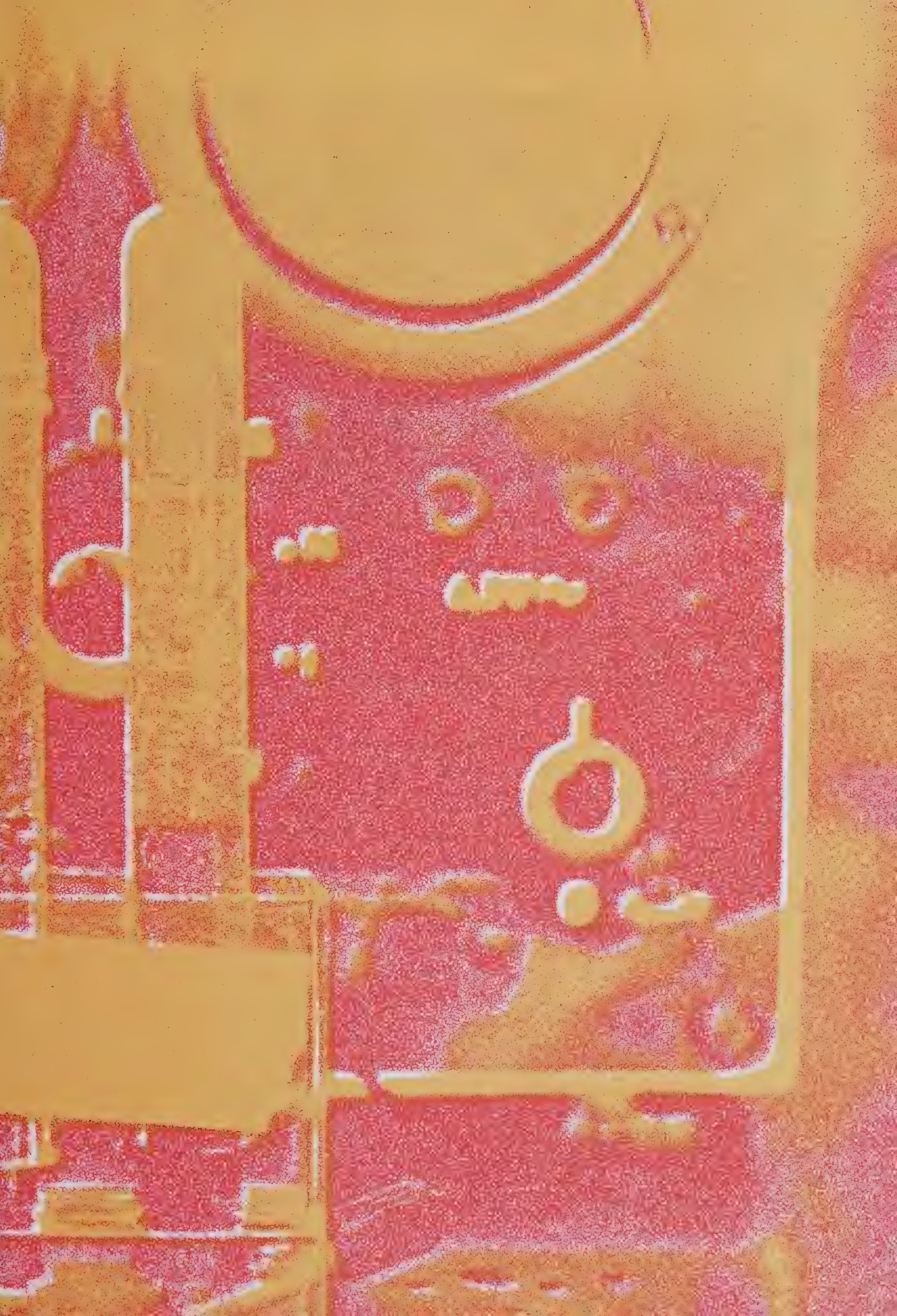
10. How many moles are there in:
    (a) 100 g of NaOH;
    (b) 416 g of C_2H_5OH;
    (c) 31.5 g of HNO_3?
11. What is the gram molecular mass of:
    (a) $CaSO_4$;
    (b) Cl_2;
    (c) NO_3?
12. A mixture of gases consists of 1 ℓ of oxygen gas, 2 ℓ of nitrogen gas and 0.5 ℓ of carbon dioxide gas. If the total pressure in the container holding the gases is 750 mm, what is the contribution to the total made by each of the gases? (All are at the same temperature and pressure.)
13. A sample of carbon with a mass of 12.0 g was burned in pure oxygen. The resulting carbon dioxide had a mass of 44.0 g. How much oxygen was consumed?
14. Use the data from Question 13 to calculate the percentage of carbon in carbon dioxide.
15. Calcium oxide is approximately 70 percent calcium. What is the percentage of oxygen in calcium oxide?
16. A sample of Rochelle salt with a mass of 28.2 g was heated to drive off the water of hydration. The anhydrous salt which remained had a mass of 21.0 g. Calculate the percentage of water in Rochelle salt.
17. A certain gas occupies 1.0 ℓ at 25°C and 760 mm pressure. What volume will it occupy at 0°C and 700 mm pressure?
18. A cylinder of compressed gas contains 20 ℓ measured at 25°C and 1200 kPa. What volume of gas will be delivered at STP?
19. (a) In an experiment to find the gram molecular mass of a gas, 1156 ml were collected at 37°C and 800 mm pressure. If the mass of this much gas is 1.34 g, what is the gram molecular mass of the gas?
    (b) Select the gas from the following list of gases that could be the one referred to in (a).
        (i) CO_2;
        (ii) CO;
        (iii) H_2;
        (iv) HCl.
20. Outline the contribution made to science by each of the following:
    (a) Joseph Gay-Lussac;
    (b) Robert Boyle;
    (c) John Dalton;
    (d) Amadeo Avogadro;
    (e) Jacques Charles.

Some Words You Should Know

density
pressure
buoyancy
diffusion
combustion
temperature
heat
element
atom
compound
mixture
molecule
water of hydration
hydrate
reagent
chemical reaction
decomposition
synthesis
pure substance
regularity
atomic mass
molecular mass
composition
equation
homogeneous
heterogeneous
Law of Conservation of Matter
Law of Definite Proportions
Boyle's Law
Charles' Law
Avogadro's Hypothesis
Kelvin
Celsius
absolute zero
kinetic energy
eudiometer
symbol
mole
STP
molar volume
Ideal Gas Law
partial pressure
vapour pressure

3

Atoms and Elements/ Molecules and Compounds/Electricity and Matter/ Fundamental Particles/ The Rutherford Atom/ The Bohr Atom/ Ions-Ionization-Ionic Bonds/Atomic Structure/Valence/ Octet Theory/ Ionic Solids/Covalent Bond/Symbolism and Formulas/Nomenclature/ Equations/Lewis Diagrams/Isotopes



Staff and research students at the Manchester Physical and Electro-Technical Laboratories, 1912. *Row 1:* R. Rossi, H. G. J. Moseley, J. N. Pring, H. R. Gerard, E. Marsden. *Row 2:* H. Geiger, W. Makower, A. Schuster, E. Rutherford, R. Beattie, H. Stansfield, E. J. Evans. *Row 3:* C. G. Darwin, A. S. Eve, P. C. H. Florence, M. White, M. S. Leslie, H. Robinson, A. S. Russell, unidentified, unidentified. *Top Row:* J. M. Nuttall, W. Kay, unidentified, J. Chadwick. (Photo courtesy of Dr. Kasimir Fajans.)

The Molecular Model

Since the time of the early Greeks men have been trying to understand the nature of matter. By the mid-1800s most scientists were fairly convinced that molecules and atoms were a good model for matter. By the mid-1900s the idea of atoms and molecules was being taught in public schools all over the world. Hopefully the previous unit gave you some reasons why the molecular model is so useful to us.

And yet no one has really *seen* an atom or a molecule. We believe in them, not because we can see them, but because they are useful to us in explaining some of the properties of matter. Our study of gases and their behaviour has provided us with the evidence we need to support this mental picture or model of gases.

We must always remember—students and teachers alike—that this picture may not be correct. Molecules and atoms might only be a figment of our imagination. Matter *may* be composed of quarks and demi-quarks!

If we decide to keep our theory of atoms and molecules for another unit some new ideas can be developed from it. For example, since we know that some molecules can be decomposed, we can deduce that some molecules are composed of simpler substances. Molecules are not necessarily the simplest form of matter. These simpler substances which cannot be decomposed are called *elements* and the particles of which elements are composed are called *atoms*.

If we apply a reverse kind of reasoning, it appears that atoms can join together to form molecules. We might ask why two atoms would join together. What is the force that binds them? Is there a limit either to the number or kind of atoms that will unite? How are the numbers determined?

We believe that all matter is made up of atoms joined together in various combinations and arrangements to form molecules. If all the atoms in a molecule are identical, we have an *element*. If, as for example in water, the atoms are different we have a *compound*.

Sulphur atoms join together to form a molecule containing eight atoms. It is an element. Hydrogen and oxygen atoms combine to form water. Water is a compound.

So far you have come across several elements: hydrogen, oxygen, magnesium, iron, iodine, carbon, nitrogen and sulphur. How does one element differ from another? How does oxygen differ from iron? Why are they different? There must be some explanation for these differences. The explanation must lie within the atom itself.

How can we look inside something so small we cannot even see the outside of it? We must obviously sharpen our other senses as we attempt to probe into the nature of the atom.

The following experiment will serve to show you what we mean.

3.1 The Black Box

Chemists believe that matter is made up of particles much smaller than anything we will ever see, even with the most powerful of microscopes. Besides believing in the existence of these particles, they have discovered quite a bit about the particles. You are probably asking yourself how anyone could know much about something that cannot be seen. To show you how such information is obtained try the following exercise.

You will be provided with a box that is sealed so that you cannot look inside. The box is not empty; it contains several objects that you will not be told about. You will not even be told how many things are in the box.

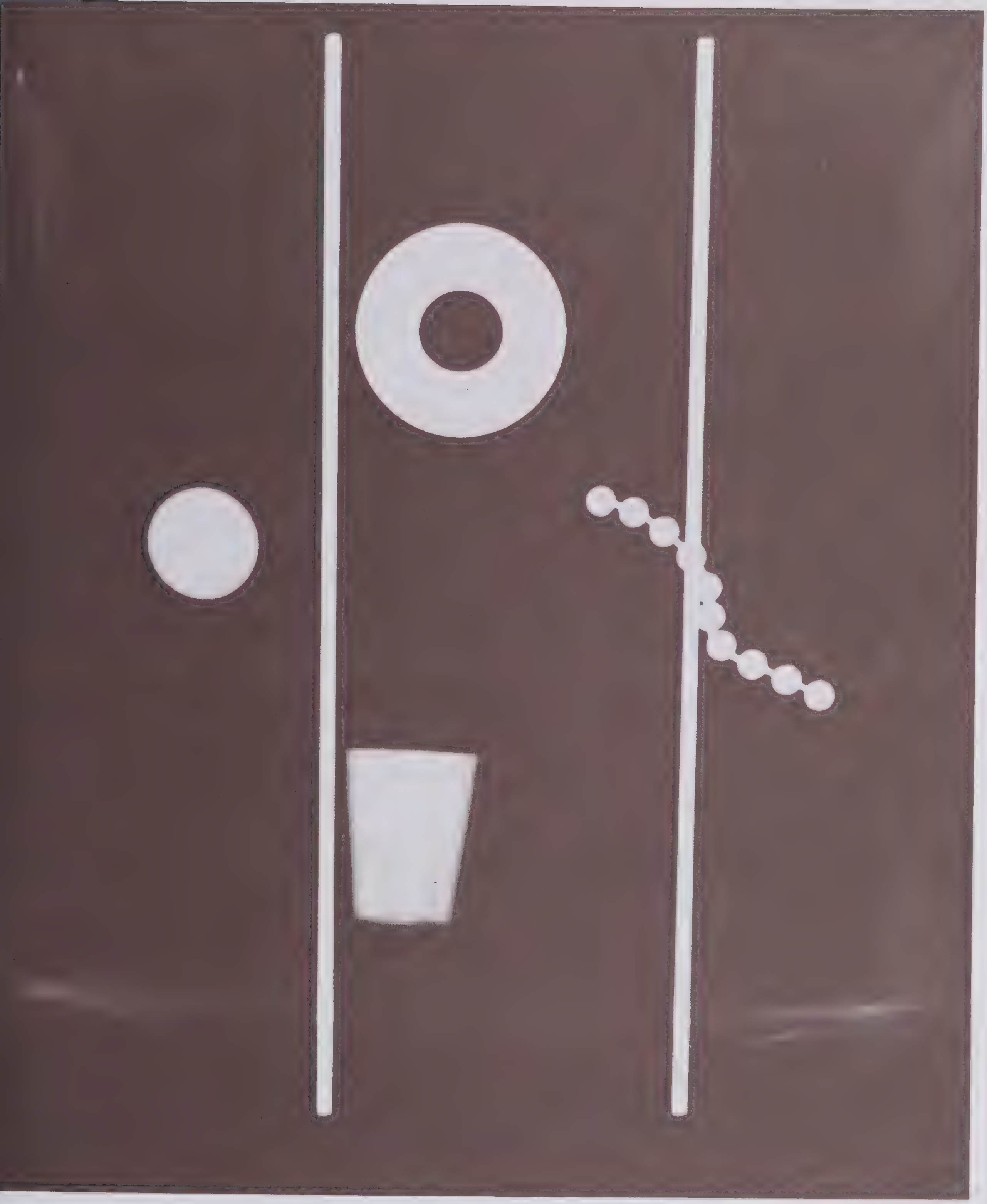
The purpose of the exercise is for you to examine the box in every possible way you can think of and learn as much as you can about the objects in the box. You must not, however, look inside the box.

Questions

1. Were any of the objects in the box round? Did any of the objects roll smoothly as you tilted the box?
2. Did the objects make a noise when they collided?
3. Were any of the objects affected by a magnet?
4. Could you tell how many objects were in the box by the sounds made as the box was moved or by the way the objects moved when the box was tilted?
5. As you gradually tilted the box did all the objects in it begin to move at the same time?
6. What questions can you ask yourself about the contents of the box?
7. What other tests do you wish you could have done with the box? (Sorry, but you still can't open it.)
8. Look at the picture on the next page. How do you think it was made? What information does it convey to you?
9. Does the picture allow you to "see" inside the box?
10. Do you think that you can describe the contents of the black box with great accuracy or only with slight accuracy?
11. What information about the contents of the box is still lacking?

A Point for Discussion

In order to continue our investigation into the nature of matter, we would like you to try and recall some work you did in physics some time ago.

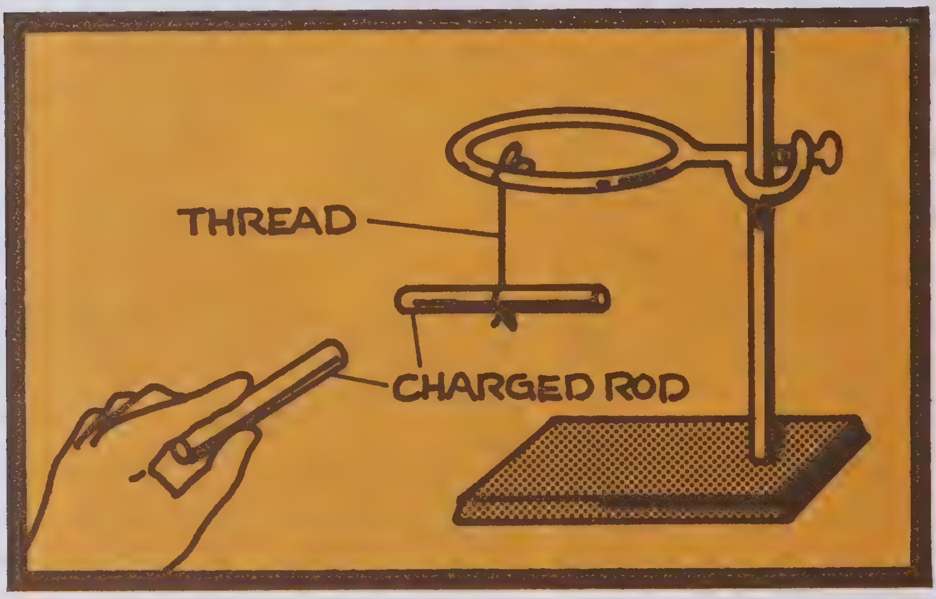


Radiograph By Bernie Ziomkiewicz Faculty of Education Queens University

Do you remember what sort of things happen if you set up an experiment such as the one shown here?

What happens if you comb your hair vigorously and then hold the comb close to some small pieces of paper? Why does this happen?

What other combinations of materials can produce a similar effect? How do we explain this strange but very common phenomenon usually referred to as static electricity?



3.2

Static Electricity (Optional Experiment or Demonstration)

ebonite rod or plastic comb
glass rod or tubing
fur
silk
thread
tissue paper

1. Bring an ebonite rod or a plastic comb near some small pieces of tissue paper.
2. Rub the rod or comb with fur (or comb your hair vigorously) and bring the rod or comb close to the bits of paper again.
3. Set the rod or comb aside for 15 min or so and proceed with the rest of the experiment. At the end of that time bring the rod or comb close to the tissue paper again.
4. Repeat Steps 1 to 3 using a piece of glass rod or tubing and rubbing it with silk.

5. Join forces with the group beside you for this next section. Carefully suspend one rod or comb by a thread from a convenient place as shown in the diagram.
6. Rub the rod or comb with fur to charge it. Bring a second rod or comb that has been rubbed with fur close to the suspended one.
7. Bring a glass rod that has been charged by rubbing with silk close to the suspended rod.
8. Bring the charged ebonite rod or comb close to a stream of water flowing freely from a water tap.
9. Bring the charged glass rod close to the stream of water.
10. Vigorously rub a plastic garbage bag with a wool sweater or jacket (or vice versa if you wish).
11. Bring a wax filled aluminum pie plate with an insulated handle in contact with the plastic bag. Touch the metal part of the pie plate with your finger. Remove the finger from the plate then remove the plastic bag from contact with the pie plate.
12. In a darkened room, bring the metallic part of the pie plate in contact with the terminals of a fluorescent light bulb. You may need to practise this last part a couple of times to make it work.

Questions

1. Why does the unrubbed ebonite or glass have no effect on the tissue paper?
2. Why does the rubbed ebonite or glass have an effect on the tissue paper?
3. How was this charge obtained? What does that tell you about matter?
4. How many different types of charge were you able to identify? What is each type called?
5. What can you say about forces between these different charges?
6. What happens to the charge when a charged object is allowed to sit for several minutes? Can you explain why this happens?
7. Would the ebonite or glass become charged just sitting on the desk?
8. What role does rubbing play in this experiment?
9. Since unrubbed ebonite and glass are not charged, what do you know about the relative amounts of electricity in these objects?
10. What does this experiment tell you about the nature of matter?
11. Can there be any doubt in your mind that matter and electricity are somehow linked?
12. Can you suggest one possible explanation for the formation of lightning in a thundercloud?

Electricity and Matter

We cannot progress any further until we go over this very important area of static electricity. There is much here of great importance to us.

We are all aware that if two different substances are rubbed together they may produce a charge. One of the best ways to get a charge is to drag your feet across a woollen rug. You'll know you were charged when you reach for the doorknob. Somehow, touching other things seems to cause the charge to diminish.

Where does the charge come from? You were not charged before you crossed the floor. The charge must have come either from you, the rug or the air around you. Furthermore, the addition of energy, in the form of all that friction, seems to be a factor.

Your experiments with static electricity would seem to indicate that there are only two types of static electricity. Further, it seems that two charges that are the same tend to repel each other. It seems, therefore, that unlike charges are responsible for the attraction you have witnessed.

This leads us to a statement of the First Law of Electrostatics. Unlike charges attract each other; like charges repel each other.

As you, the rug and the air are all composed of matter, it seems reasonable to suppose that matter and electricity have something in common. We must not forget that energy is also involved.

If you carry out some more simple experiments you will discover that when two different objects are rubbed together and a static buildup is observed, *both* the objects become charged and the charges are opposite.

Now it is time for some terminolgy. Let us first label the two types of charge. For convenience we will call one *positive* and one *negative*. We can only tell them apart by comparing them with a known charge. Somewhere along the line we must define something. The oldest definition was put forward about 1800. The charge on an ebonite rod that had been rubbed with cat's fur was defined as negative. That means any charge attracted to ebonite must be positive and any charge repelled by ebonite must be negative. There have been more recent definitions but none have been any simpler than this one.

We should now go back to the question, why does the charge build up? What is happening when two substances are rubbed together?

Let us begin with what we know. Most substances have no charge most of the time. Such substances are said to be *neutral*. When two neutral substances are rubbed together, one sometimes becomes negatively charged and the other becomes positively charged. But where does the charge come from? There are at least three possibilities:

1. from within the positively charged substance;
2. from within the negatively charged substance;
3. from some external source.

Static charge can build up in a vacuum, which would seem to eliminate the third possibility. That leaves us with the idea that charge comes from within a substance.

If some substances have charges within them, why would they appear to be neutral most of the time?

If a substance had equal amounts of both positive and negative charge, then it would appear to be neutral. If, somehow, one kind of charge or the other could be "rubbed off," then both the "rubb-er" and the "rubb-ee" would appear to become charged. Here is an example.

Suppose you have 1000 positives and 1000 negatives in you. Suppose a wool rug has 5000 positives and 5000 negatives. Both you and the rug appear to be neutral. Now suppose you drag your feet across the rug and 100 units of negative are transferred by this energy input from the rug to you.

You now have	1100	negative
	1000	positive
net	100	negative
And the rug has	5000	positive
	4900	negative
net	100	positive

You appear to be negative and the rug appears to be positive. Both are charged with opposite charges, and that is consistent with known fact.

We have then produced a new model for matter that explains static charge buildup. The central points of our model or theory are:

1. Matter contains both positive and negative electricity.
2. A neutral object contains equal amounts of positive and negative electricity.
3. Charge can be transferred from one substance to another if energy is provided.
4. Like charges repel; unlike charges attract.
5. There is a relationship between matter, electricity and energy.

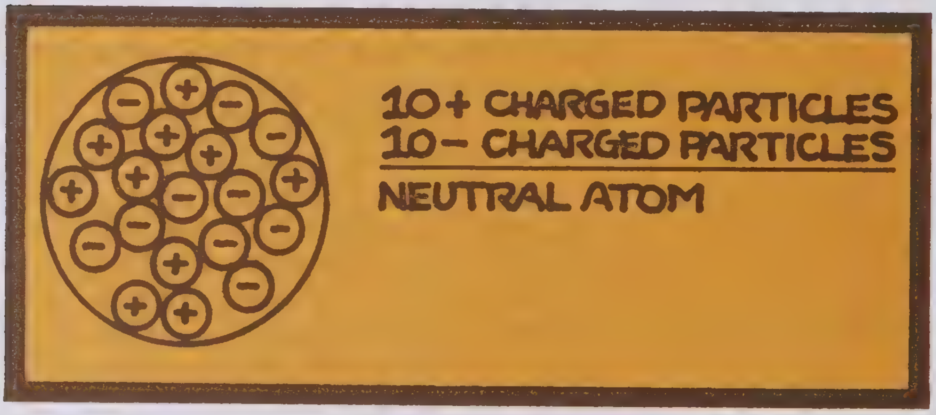
What do we know so far? Matter is composed of atoms that contain electrical charges. From our electrostatics we might conclude that energy (friction) may cause charge to be transferred from one atom to another. However, the neutral atom must contain equal amounts of positive and negative charge. Let's try to draw the atom or at least imagine what it looks like. Of course we cannot see charge but let's go ahead anyway.

Imagine that the positive charge looks like a peanut. Count out 20 of them (or any other number you wish). Imagine that the negative charge looks like a raisin. Count out 20 of them (or the same number as you did peanuts). Now mix them together by shaking them in your cupped hands. Form the mixture into a ball shape and "Presto"! If you remove your hands it falls apart. Atoms don't do that. Our model must lack something. (Actually some atoms do fall apart as you will see later.)

Let's revise our picture a little. Imagine that the positive charge looks like pancake batter. One spoonful of batter is equivalent to one unit of charge. Measure out 20 spoonfuls of batter. Add 20 raisins (20 negative charges), mix well and cook over a hot griddle.

After cooking this model will stay together. Twenty positive and 20 negative make it neutral. It's also flat and somehow that doesn't seem right, but we'll leave you to figure out how to cook a spherical pancake!

Why spherical? Somehow that just seems neater in our imagination. At the moment you are entitled to any shape you wish as long as it is neutral.



The Thomson Atom

These ideas about the arrangement of charges in the atom were first proposed by an English physicist, J. J. Thomson. About the only difference is that J. J. (as he was known to his friends) used plums instead of raisins and dark cake instead of pancake. His recipe produced the traditional English Christmas pudding or plum pudding. It's not surprising then that a rather sceptical press christened J. J.'s model, "The Plum Pudding Atom."

His was the first model that tried to incorporate the electrical nature of matter. Developments and modifications to the model came thick and fast. Work was going on in many laboratories in many countries. Some say that Thomson's model of the atom marked the beginning of the great Age of Science which began about 1900 and continues unabated today.

We must digress for a moment. Many men had done work in the area of matter and electricity years before Thomson's model.

In about 1780 Alessandro Volta had discovered the voltaic pile.

J.J. Thomson

This was a device capable of producing electricity that is today called a battery.

John Grey in 1790 had shown that static charge could be conducted along a wire and gave us current electricity.

We are not trying to burden you with names and dates, but we do feel it is important for you to recognize that the achievements of one man were often supported by the work of many others before him.

Sprengel had invented the vacuum pump and this lead the way for Plucker to do experiments using Volta's batteries. He found that low-pressure gases conducted electricity. But it was a man named William Crookes who took the experiment and refined it. It was also Crookes who tried to find the meaning of it.

Crookes' apparatus consisted of a glass tube sealed at both ends. Inside the tube were two electrodes and an opening connected to a vacuum pump. This is known as a Crookes' tube.

3.3

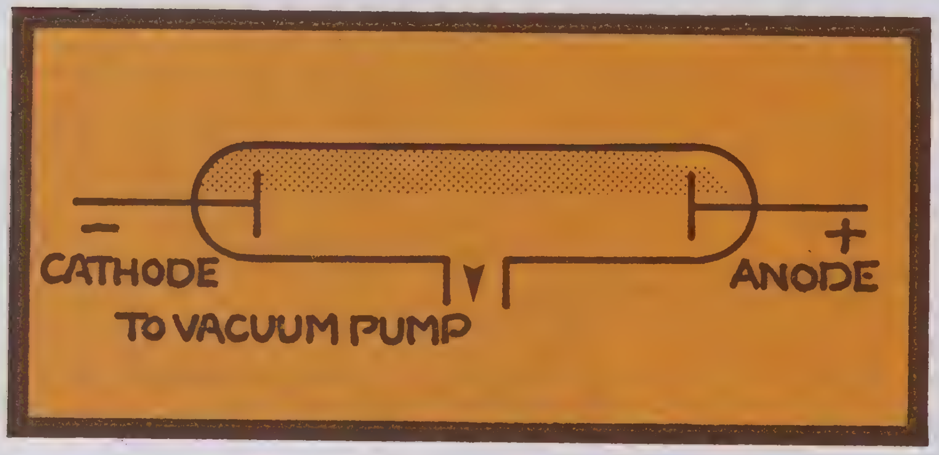
Cathode Ray Tubes (Demonstration)

Warning

Use the lowest possible voltages in this experiment to avoid producing unnecessary radiation (x-rays).

Your teacher will demonstrate for you some or all of the following pieces of equipment. Each demonstration will add one more piece of information that we may include in our electrical model for matter. If you do not have this equipment, please accept our observations.

The first demonstration will involve a piece of equipment called a Crookes' tube. The set up will look like the diagram below.

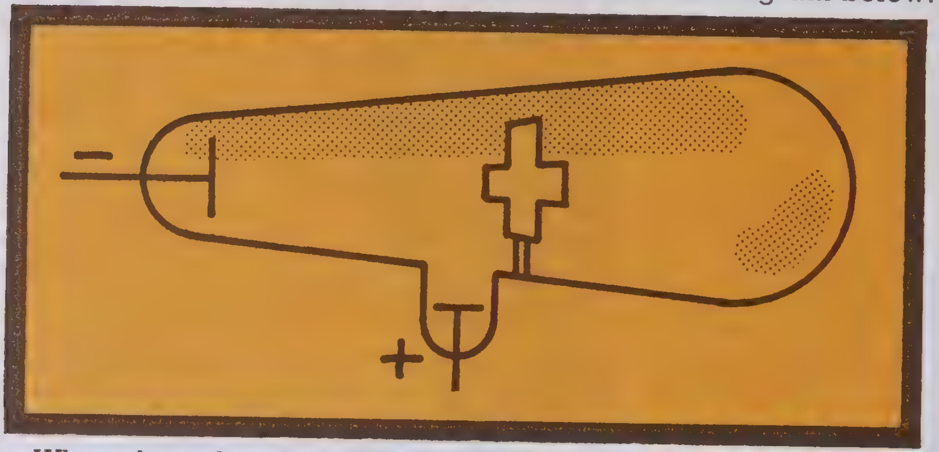


The electrodes are first attached to a high voltage source. The room lights are turned off and the power switched on. Then the pressure inside the tube is reduced with a vacuum pump. The pink rays observed inside the tube at high voltages and low pressures are called *cathode rays*. Cathode rays may be true waves, that is, energy is transmitted from one end of the tube to the other, or they may be a stream of particles moving from the cathode to the anode. In any case we know that these rays originate at the cathode (matter) under the influence of electricity. Again we have a link between matter and electricity.

One more thing — if the rays are made of particles then the particles must carry a negative charge since they move away from the negative cathode and towards the positive anode.

Sir William Crookes

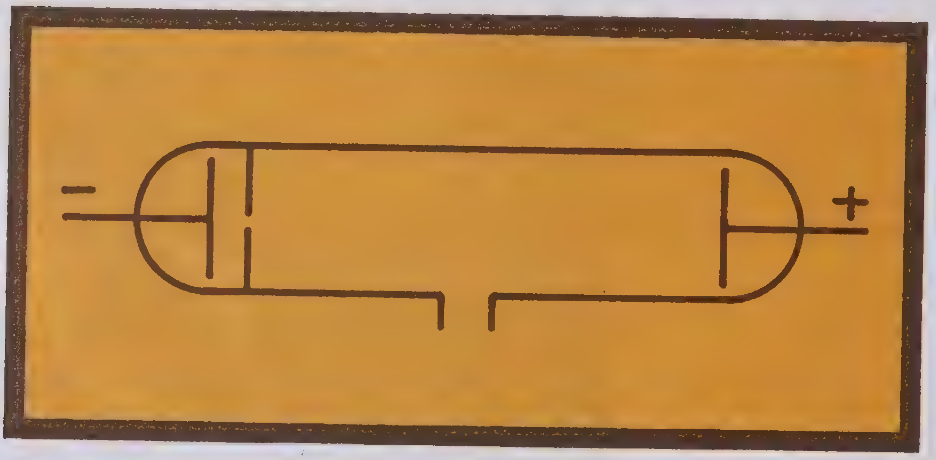
The second type of tube to consider is like the diagram below.



When the voltage is applied several things happen. Cathode rays appear moving down the tube. The shadow of the cross appears on the end of the tube and the glass around the outside of the shadow fluoresces or glows.

This tube tells us that cathode rays tend to travel in straight lines otherwise there would be no clear shadow. Secondly the ray cannot penetrate the material the cross is made from and thirdly, when they strike the glass they cause it to glow. The latter seems to indicate that cathode rays contain energy.

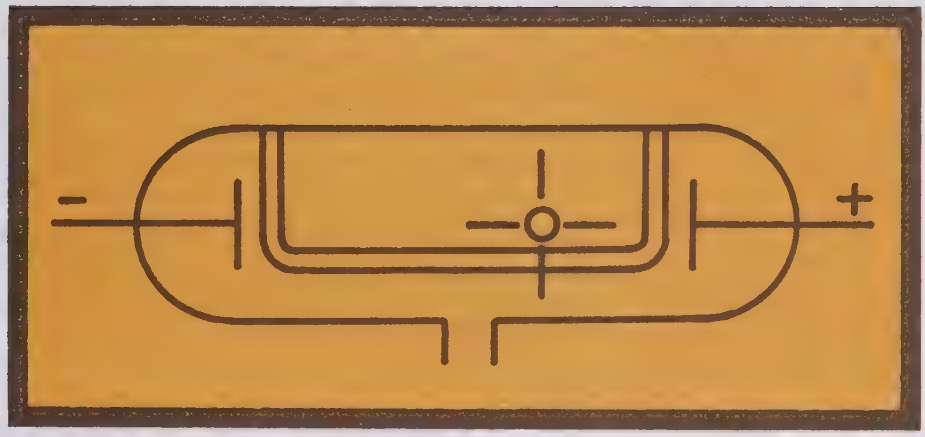
The third type of tube to consider is like this diagram.



In this tube we see the slit produce a beam that travels straight down the tube. If a magnet is moved near the beam, the beam is deflected and follows a curved path where it is near the magnet. One pole of the magnet attracts the beam, the other seems to repel the beam.

We know that waves such as sound and light cannot be deflected by a magnet so perhaps our cathode rays are not waves of energy. That leaves the idea of a stream of particles — a stream of negatively charged particles — as the explanation.

One more tube can help give us an answer or confirm our suspicion.



When this tube is turned on the tiny paddlewheel moves from the negative terminal towards the positive terminal. If the polarity is reversed then the direction of travel of the paddlewheel is also reversed. Apparently some energetic particle is striking the vane of the paddlewheel and is causing it to turn just as a waterwheel or windmill is turned.

J. J. Thomson looked at the evidence you have just seen and decided that cathode rays were composed of a stream of negative particles being repelled from the negative cathode and attracted to the positive anode. He even went so far as to give his proposed particle a name — the *electron*.

To summarize then:

1. Cathode rays were produced when high voltages were used in conjunction with low pressure.
2. Cathode rays tend to travel in straight lines.
3. Cathode rays are not very penetrating.
4. Cathode rays cause certain types of materials to fluoresce.
5. Cathode rays are deflected by a magnetic field.
6. Cathode rays have kinetic energy and therefore mass.
7. Cathode rays carry a negative charge.

Sub-Atomic Particles

Thomson's research continued until he was able to estimate the mass of the electron as about 1/2000 the mass of the lightest atom known — hydrogen. Here then was the proof that the atom was composed of even smaller *sub-atomic particles*. How well this fits into the model of the plum pudding described earlier! How easy it was then for Thomson to look into the work of one Eugene Goldstein who had discovered a ray he called a canal ray and which went in the opposite direction to a cathode ray. Thomson decided that these canal rays must represent a beam of positive particles. Wilhelm Wein actually showed the positive character of the ray but it was Thomson who determined that their mass was about equal to the mass of a hydrogen atom. Thomson called them *protons*.

Hard work by several people including Thomson, Goldstein, Wein and Millikan showed that the proton and the electron had the same amount of charge; it was the quality of the charge that was different. We can summarize what we know in tabular form.

Table 3.1

Particle	*Mass*	*Charge*	*Symbol*
hydrogen atom	1 u*	neutral	H
electron	1/2000 u	negative (−1)	e−
proton	1 u	positive (+1)	p+

*u — atomic mass unit

Since atoms are neutral they must be composed of equal numbers of protons and electrons. Our atom is now composed of charged particles called electrons and protons packed together in equal numbers to form a small dense package. J.J. Thomson's Plum Pudding Atom was the first serious explanation of the electrical nature of the atom.

Sir William Crookes had started it all in about 1860 when he discovered cathode rays. The model was complete in about 1910 when Millikan made his accurate measurements of the charge on the electron as 1.6×10^{-19} C*. It took about 50 years to build up the necessary knowledge to build the model. It took just one experiment to break it down.

3.4 Probing the Unknown

plasticine
pin

1. Your teacher will give you a lump of plasticine that has a metallic object embedded in it.
2. Using only a pin, you must try to determine the shape and dimensions of the metallic object.
3. There are three ground rules to the exercise.
    1. The pin may not be wriggled around inside the plasticine.

---

*C = symbol for coulomb. One coulomb is the charge transported in one second by a current of one ampere (1 C = 1 A·s).

2. Each pin hole must be covered before the next jab is made.
3. You may not change the shape of the lump of plasticine.

Questions

1. How did you use the pin to solve the problem?
2. Why does this technique work in this experiment?
3. How many tries did it take to establish the shape of the metal and its location roughly? Accurately?
4. Could you do this experiment this way if the piece of metal were smaller than the pointed end of the pin? Why?
5. What sort of pin would you have to use if you wanted to find a single atom?

Other Discoveries

As we work our way through the sciences we often find references to discoveries that were made by "accident." When we consider the number of accidental discoveries we cannot help but wonder whether they were, in fact, accidents at all. Dr. Hans Selye, a director of medical research at the University of Montreal, would probably say they were not. He would probably add that they were the result of observations made by trained observers, observations that untrained people would either have missed or worse still, dismissed as not being relevant to the experiment at hand. In any case, three discoveries that some have called accidental pushed science ahead very rapidly at this point.

In 1896 Henry Becquerel, a French physicist, accidentally discovered that a lump of the uranium ore called pitchblende fogged or clouded unexposed photographic film. It was as though the ore produced something that had the same effect on film as exposure to light. Other scientists soon found that if a lump of pitchblende were brought near an electrostatically charged object, the charge was lost much faster than usual.

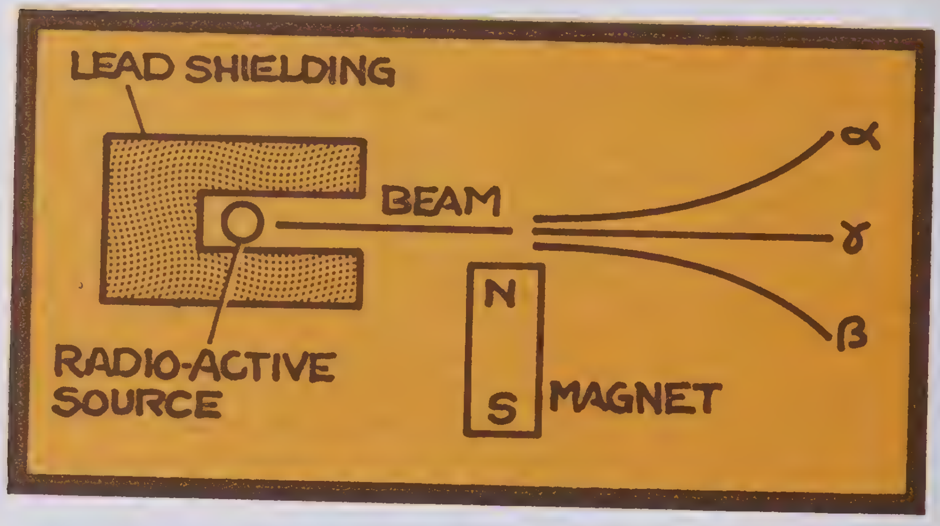
Pierre and Marie Curie discovered that the "thing" responsible for these two phenomena emanated from the elements uranium and radium. In the process they also discovered a brand new element which they named polonium after Madame Curie's native land. Work with the *emanations* from these three elements soon showed that they were similar in many ways to the very penetrating and mysterious rays that William Röentgen had discovered — by accident — in 1895 and named x-rays.

3.5 The Geiger Counter (Demonstration)

At this point if your teacher has a radioactive source and a geiger counter he can demonstrate the penetrating capacity of these x-ray-like emanations.

The Rutherford Atom

The new substances were labelled "radioactive" by our good friend J.J. Thomson and his young New Zealand research companion, Ernest Rutherford. Their work with these substances soon led to another discovery. If a beam of radioactive rays was passed near the pole of a magnet the beam was split into three distinct parts. The parts were labelled alpha, beta and gamma after the first three letters of the Greek alphabet.



The alpha beam was attracted by a negative electrostatic charge. It was not very penetrating; in fact a sheet of paper or a few centimetres of air stopped the alpha rays cold.

The beta beam was attracted by a positive electrostatic field. It was only a bit more penetrating than the alpha beam. More research later showed that the beta beam was identical in many ways to cathode rays. It was, therefore, considered to be a stream of fast moving electrons.

The gamma part of the beam was found to be very penetrating. They could, like x-rays, penetrate wood, glass, metals, stone, concrete and flesh. X-rays of course had been tamed and were being used to take pictures inside the body. It seemed that about the only difference between gamma and x-rays was that gamma rays were more energetic and penetrating.

Ernest Rutherford left England in 1898 to take up a teaching and research post at McGill University in Montreal. Here, in the heartland of French Canada, Rutherford and his new research team made several more important discoveries.

First they discovered, as you might have already suspected, that the alpha ray was not a ray but a beam of particles. Compared to the electron they were great heavy particles and carried a positive charge. It was determined that they had about the same mass as a helium atom and had double the charge of a proton. Such a charged particle was called an ion — alpha particles were helium ions: He^{++}

The team's second discovery was that when the atoms of a radioactive substance give off their emanations of alpha, beta and gamma radiation they change into a different element. It was possible, as the alchemists had guessed 400 years earlier, for one element to change into another.

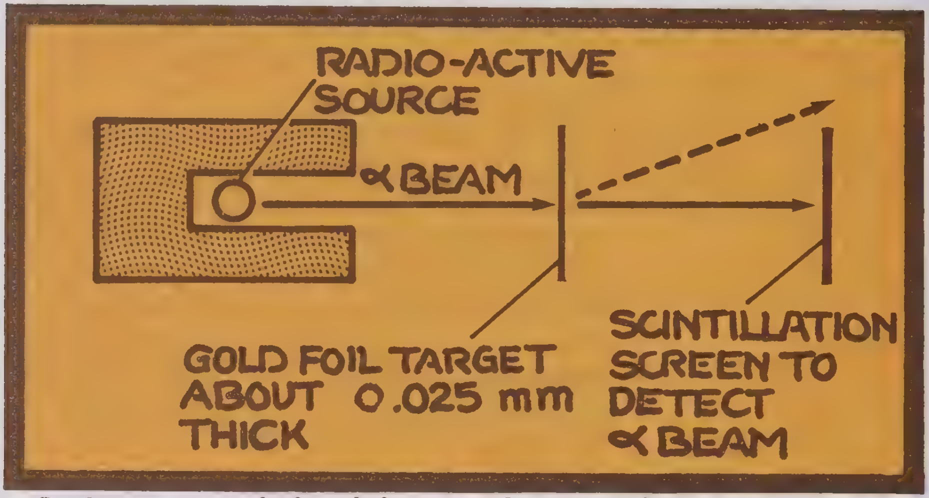
$$\text{Polonium} \longrightarrow \text{Lead} + \alpha + \beta + \gamma$$
$$\text{Po} \longrightarrow \text{Pb} + \text{He}^{++} + \beta + \gamma$$
$$\text{Uranium} \longrightarrow \text{Thorium} + \alpha + \beta + \gamma$$
$$\text{U} \longrightarrow \text{Th} + \text{He}^{++} + \beta + \gamma$$

The stage was set for the "accident" that was to make three men famous.

Rutherford returned to England in 1907 to become professor of physics at Manchester and it was here that he accepted Hans Geiger as his laboratory research assistant. Together they perfected a device for "seeing" an alpha particle. It was a metal disk coated with zinc sulphide. If an alpha particle struck the coating it gave off a burst of light called a scintillation which was visible through a microscope in a darkened room.

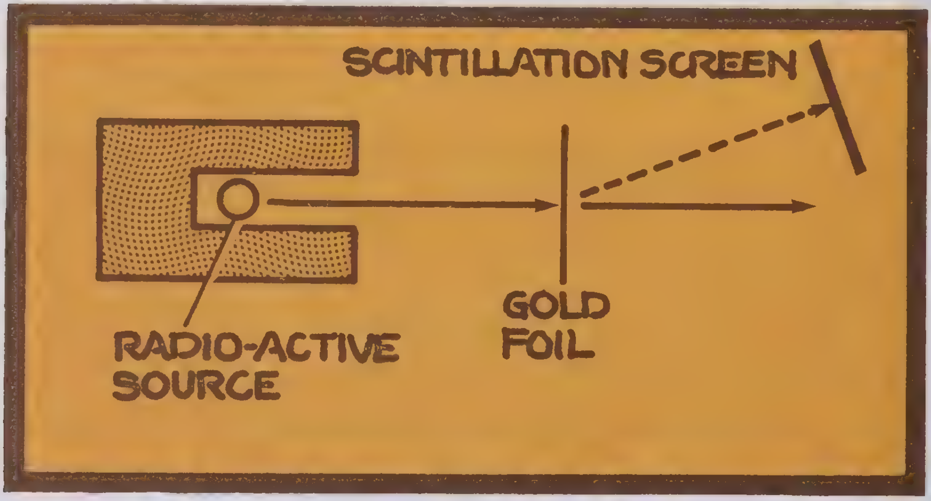
With work progressing well, Rutherford found it necessary to enlarge the team and hired a student named Ernest Marsden. Rutherford, as director of research, proposed an experiment that Geiger and Marsden were to carry out for him. The plan was simple. They were to attempt to discover what happened when a beam of alpha particles was directed at very thin samples of a metal. Rutherford was fairly sure they would sail on through because the alpha particle was so large and heavy, but the experiment would break in the new assistant and mould him into the team.

First they set up the apparatus like this diagram.

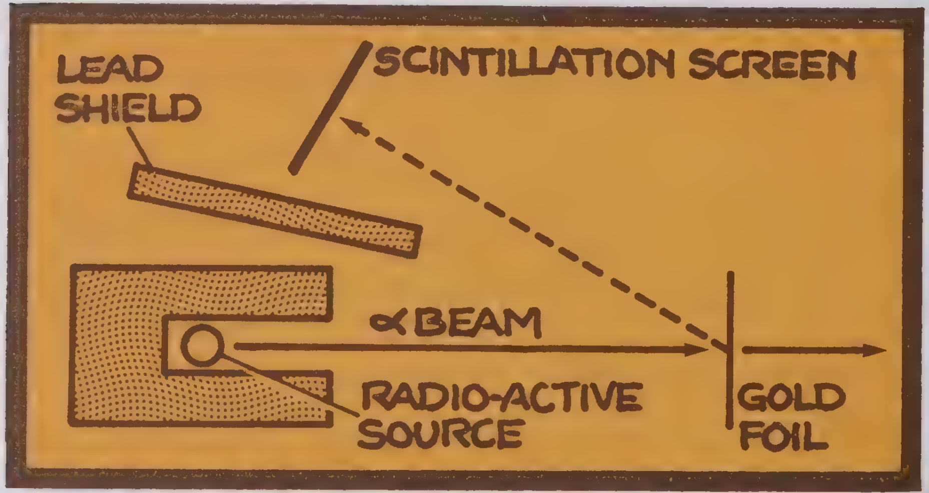

RADIO-ACTIVE SOURCE
∝ BEAM
GOLD FOIL TARGET ABOUT 0.025 mm THICK
SCINTILLATION SCREEN TO DETECT ∝ BEAM

As they expected, the alpha particles passed through although a very few seemed to be bent on the way.

The apparatus was then changed to try and discover how many particles were bent and by how much.


SCINTILLATION SCREEN
GOLD FOIL
RADIO-ACTIVE SOURCE

Finally the apparatus was set up as in this diagram.


LEAD SHIELD
SCINTILLATION SCREEN
∝ BEAM
RADIO-ACTIVE SOURCE
GOLD FOIL

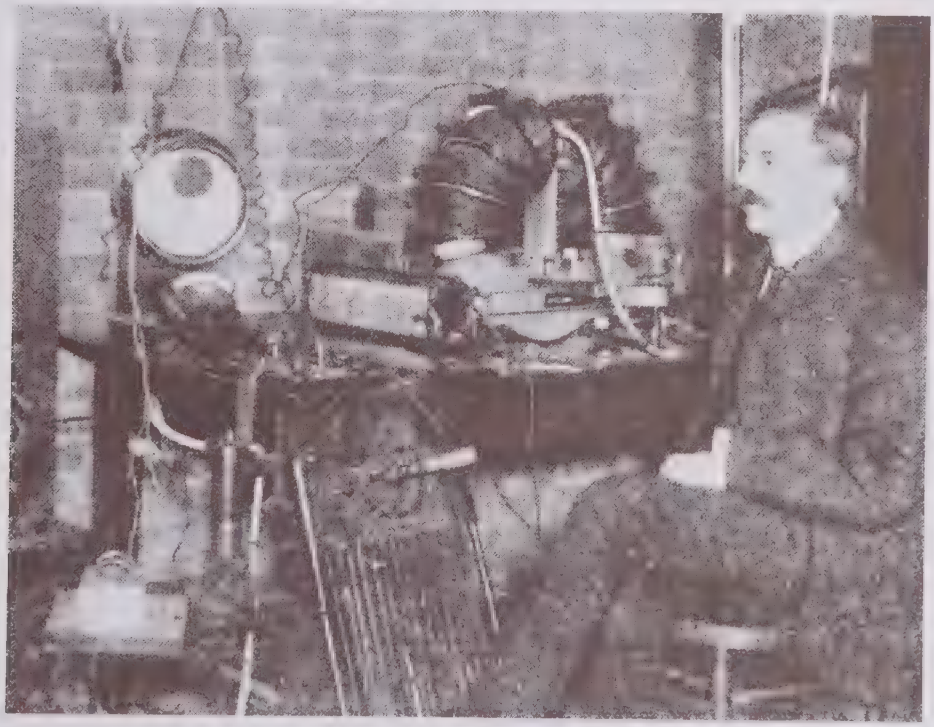
The experiment designed to break in the new assistant had startling results. Alpha particles were detected even in the final position. There were not very many but there were definitely some. They must have travelled from the source to the foil and back to the detector. When Rutherford heard of this he was astounded. In his memoirs he says of the incident, "It was as though we had fired a 15-inch naval gun at a piece of tissue paper and had the shell bounce back at us!"

Careful analysis of the data provided by the experiment gave some interesting results.

Firstly, Rutherford decided that the alpha particles were not actually hitting anything. They were being repelled by a powerful electrostatic field. Since the alpha particle was positive this meant there had to be a strong positive charge in the atom. But where?

Thomson's atom showed the positives and negatives are mixed together. Rutherford decided this could not be right. The positive charges must be in a separate place which he described as a small dense core or *nucleus*. It was this nucleus that deflected the alpha particles.

The electrons were somewhere outside the nucleus. Rutherford reasoned that if the electrons were outside the nucleus then they would be very strongly attracted to it. They should, by all that seemed reasonable, fall into the nucleus. The electrons would fall into the nucleus unless they were moving around the nucleus at a high speed in the same way as the earth and other planets orbit the



Ernest Rutherford. This Rube Goldberg machine helped Rutherford discover alpha and beta rays in 1905 at his laboratory at McGill University.

sun. Under these conditions the very strong attractive forces between the nucleus and the electrons would become a centripetal force that held the electrons in an orbit around the nucleus.

If you remember some physics:

Centripetal force

$$E_{(cf)} = \frac{mV^2}{r}$$

where m = mass
V = velocity
r = radius of orbit

The force of attraction between two charged objects is given by Coulomb's Law.

Electrostatic force

$$F_{(es)} = \frac{k_{(c)}q_1q_2}{d^2}$$

where $k_{(c)}$ = constant

$$\left.\begin{matrix} q_1 \\ q_2 \end{matrix}\right\} = \text{the two charges}$$

d = distance between charges

For an electron in orbit these two forces are balanced and distance (d) is the same as radius (r), $d = r$.

$$F_{(cf)} = F_{(es)}$$

$$\frac{mV^2}{r} = \frac{k_{(c)}q_1q_2}{r^2}$$

$$mV^2 = \frac{k_{(c)}q_1q_2}{r}$$

The mass of the electron is a constant and so is the charge on the electron. Since the mass of an electron and its charge are both constant:

$\therefore m = k_1$
$q_1 = k_2$

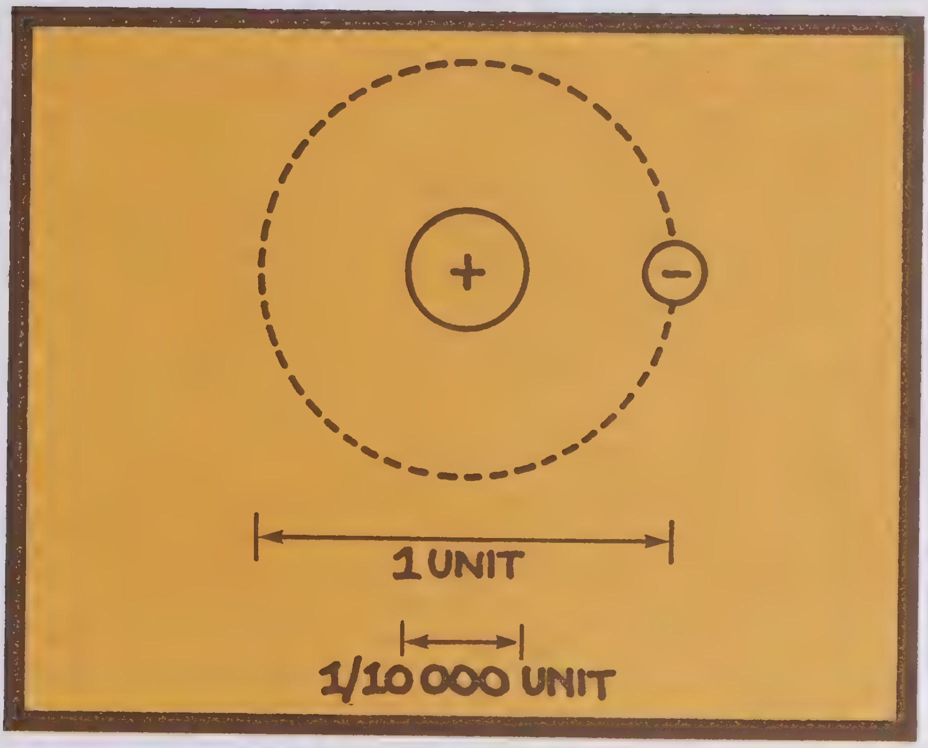
If we combine these two constants with $k_{(c)}$

$$V^2 = \frac{k_{(c)}q_2}{r}$$

$$r \alpha \frac{q_2}{V^2}$$

The radius gives us some idea of the outside of the atom. Three important points come out of all this:

1. It is not important for you to know the derivation of these equations.
2. Physics is very important to a chemist.
3. Rutherford was able to show from the results of the experiment and from his calculation, that the radius of the atom was about 5000 times the size of the nucleus. It's probably easier to talk in terms of the diameter.



It is hard to understand all this from the numbers, so try to imagine the nucleus as a ping pong ball sitting between the goal posts of a football field. The electron, however big it is, is whirling around, not one, but two full football fields away.

This is the essence of Rutherford's new model for matter. It is, of course, called the Rutherford atom. Here are the important features.

1. All the positive charge is concentrated in a small dense core called the *nucleus*.
2. The *electrons* move in orbits around the nucleus at a relatively great distance from it.
3. The electrons are held in their orbits by the electrostatic force of attraction from the nucleus.

The last point we are going to make might seem obvious. There must be a great deal of empty space inside the atom. It is not surprising that so many alpha particles passed through the gold foil undeflected.

The Rutherford model of the atom with its dense nucleus and *planetary electrons* replaced the Thomson model because it explained some of the facts better. Rutherford's atom explained the results of his alpha scattering experiment. Thomson's model did not.

So it goes in science. A model lasts until it is no longer useful, then it yields to a better model. The year was 1911. Thomson's model had lasted about 15 years. Rutherford's model would last for only two. More important perhaps, Rutherford himself recognized that his model had flaws in it!

3.6 Atomic Dimensions

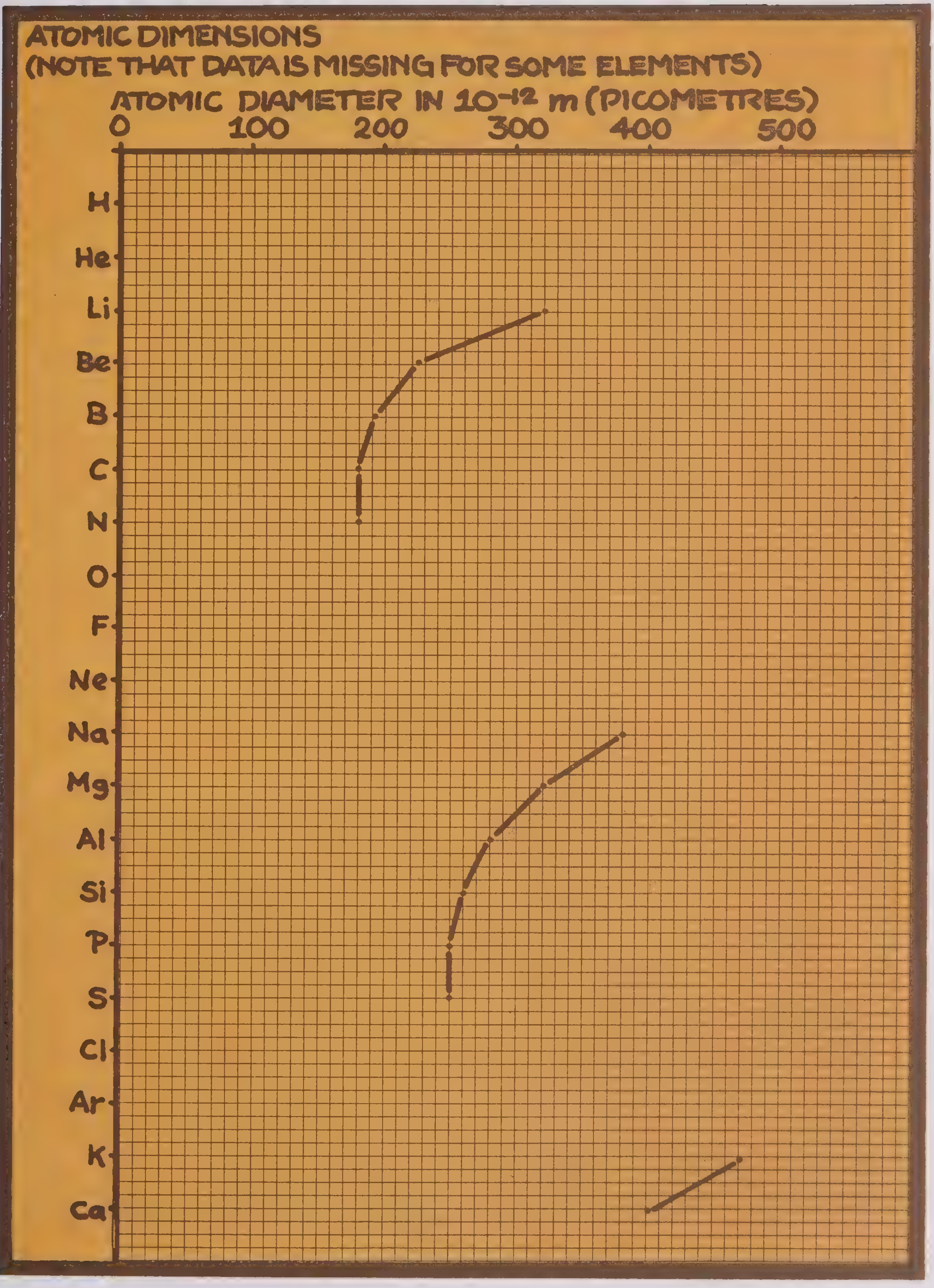
Before we go too much further we had better get some idea of just how big these atoms we are dealing with really are. The following experiment may help you to find the dimensions of an atom.

samples of various elements

1. Find the mass of the piece of lead provided by your teacher.
2. Determine the volume of the lead.
3. Calculate the density of the lead in g/cm^3.
4. Assume that you were given a piece of lead 1 cm × 1 cm × 1 cm.
   (a) What would its mass be?
   (b) How many lead atoms are present in the sample?
   (c) How many lead atoms are lined up along one edge?
   (d) What distance is there between the centres of two adjacent atoms?
   (e) How big is a lead atom?
5. Repeat this calculation for one other substance from this list.
   (a) copper
   (b) silver
   (c) iron
   (d) carbon
   (e) zinc
   (f) aluminum
6. Record the results of your calculation on the blackboard.

Questions

1. What assumption must you make to answer Question 4 (e)?
2. How do the relative sizes of atoms compare with your idea of what they might have been?
3. What is your estimate of:
   (a) The smallest object visible to the naked eye?
   (b) The smallest object visible through a microscope?
4. Look at the graph on the next page. This represents atomic diameters for the first 20 elements. Can you find any pattern to the line?



ATOMIC DIMENSIONS
(NOTE THAT DATA IS MISSING FOR SOME ELEMENTS)

Table 3.2

Element	*Gram Atomic Mass*	*Density in Solid State* g/cm^3	*Atomic Radius* $\times 10^{-8}$ cm
Hydrogen	1.0	0.10	
Helium	4.0	0.12	
Lithium	6.9	0.53	1.55
Beryllium	9.0	1.35	1.12
Boron	10.8	2.34	0.98
Carbon	12.0	2.26	0.91
Nitrogen	14.0	0.74	0.92
Oxygen	16.0	1.01	
Fluorine	19.0	0.99	
Neon	20.2	1.10	
Sodium	23.0	0.97	1.90
Magnesium	24.3	1.74	1.60
Aluminum	27.0	2.70	1.40
Silicon	28.0	2.33	1.32
Phosphorus	31.0	1.82	1.28
Sulphur	32.1	2.07	1.27
Chlorine	35.5	1.56	
Argon	39.9	1.40	
Potassium	39.1	0.86	2.35
Calcium	40.0	1.55	1.97

The Bohr Atom

The flaw in Rutherford's atom was, in simplest terms, this: when electrons move they produce a magnetic field. This requires energy, so the electrons that are moving to produce the magnetic field must give up some of their energy to the magnetic "wave." You have probably heard such energy referred to as electromagnetic energy. It does not necessarily take the form of magnetic energy as you will soon see.

If the electrons in Rutherford's atom gradually lost energy to the magnetic field then they should gradually slow down and eventually fall into the nucleus. That would put us back to a plum pudding state again! Calculations showed that the falling process should take about 10^{-10} s from start to finish. Obviously, in the billions of years the earth has been around there has been sufficient time for this to happen — and it has not!

Along came a fresh face from an as yet unheard country — Niels Bohr of Denmark. Bohr's solution to the problem of radiating energy was both dramatic and terribly simple. Bohr said what everyone knew and had left unsaid, "Electrons located in atomic orbits do *not* radiate energy unless . . . ," and he added a few qualifiers. Bohr's idea had mathematical support provided by a curious theory called the quantum theory. Your teacher may decide to tell you about quanta and energy and the man who proposed the theory, Max Planck.

Bohr said that an electron could circle the nucleus forever providing it limited itself to certain *specific* orbits. Energy was involved if the electron changed from one orbit to another but as long as it remained in a specific orbit it did not gain or lose energy.

Perhaps the most familiar energy analogy to Bohr's description of electron behaviour is a staircase. Let us suppose that each step is 0.1 m high and we have a mass of exactly 50 kg. If we go to climb the staircase we must expend energy at the rate of:

$$\text{work} = \text{force} \times \text{distance} = 50\ \text{kg} \times 0.1\ \text{m} \times 9.8\,\frac{\text{N}}{\text{kg}} = 49.0\ \text{Joules}$$

for every step we climb. At the same time we are gaining exactly the same amount of potential energy. If we were to climb two steps at once the work would be:

$$\text{work} = \text{force} \times \text{distance} = 50\ \text{kg} \times 0.1\ \text{m} \times 9.8\,\frac{\text{N}}{\text{kg}} \times 2 = 98.0\ \text{Joules}$$

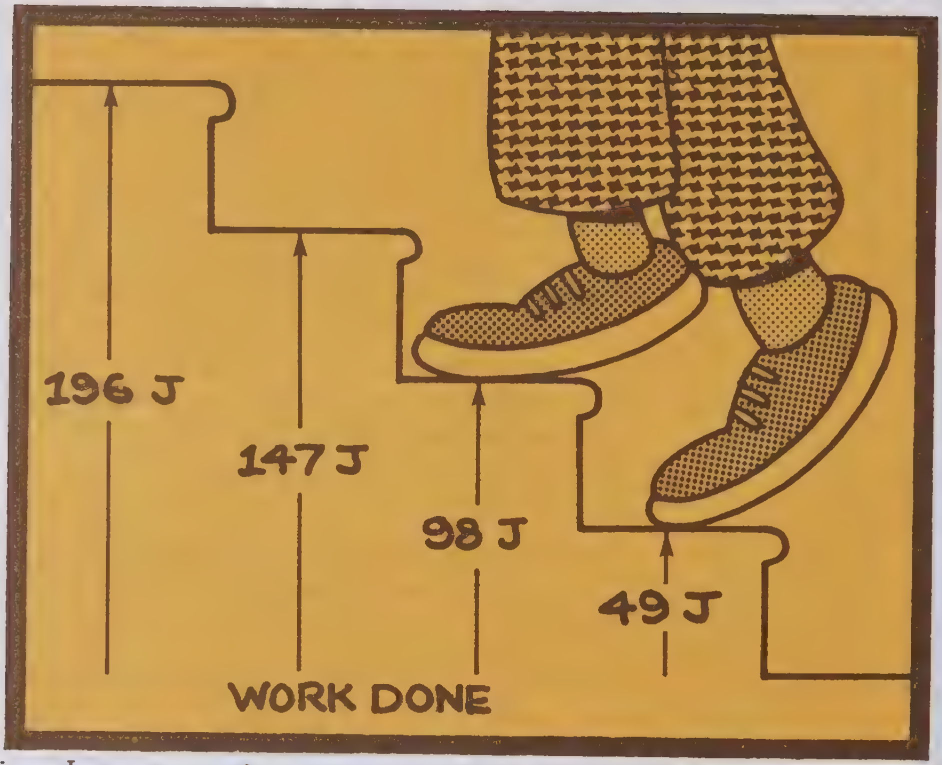
and we would gain 98.0 J of potential energy.

Here is the point of the analogy—when climbing these stairs we may gain 49 J or 98 J or 147 J (and so on) of potential energy but never 25 J or 57 J or 110 J or any amount not on the first list. There are only certain places to put your foot and therefore only certain energies are possible.

This staircase analogy is useful in considering the energy levels of an electron. It does, however, have one failing worth mention-

Max Planck

Niels Bohr



ing. In our experience, in any given set of stairs, the steps are usually all the same height—about 20 cm or so. In the atomic world the steps are all a different height.

In any atom no two steps are the same height. In fact, if we compare different atoms we find that no two elements have the same staircase. That may sound very complicated but, in fact, it is an extremely useful difference among elements.

3.7 The Atomic Staircase

So often in chemistry we refer back to physics. Remember this? "Energy cannot be created or destroyed, but can be changed from one form to another."

Let's look at the atomic staircase. All you need to remember is that light is a form of energy and that different colours (wave lengths) have different energies. The blue end of the spectrum has more energy than the red end.

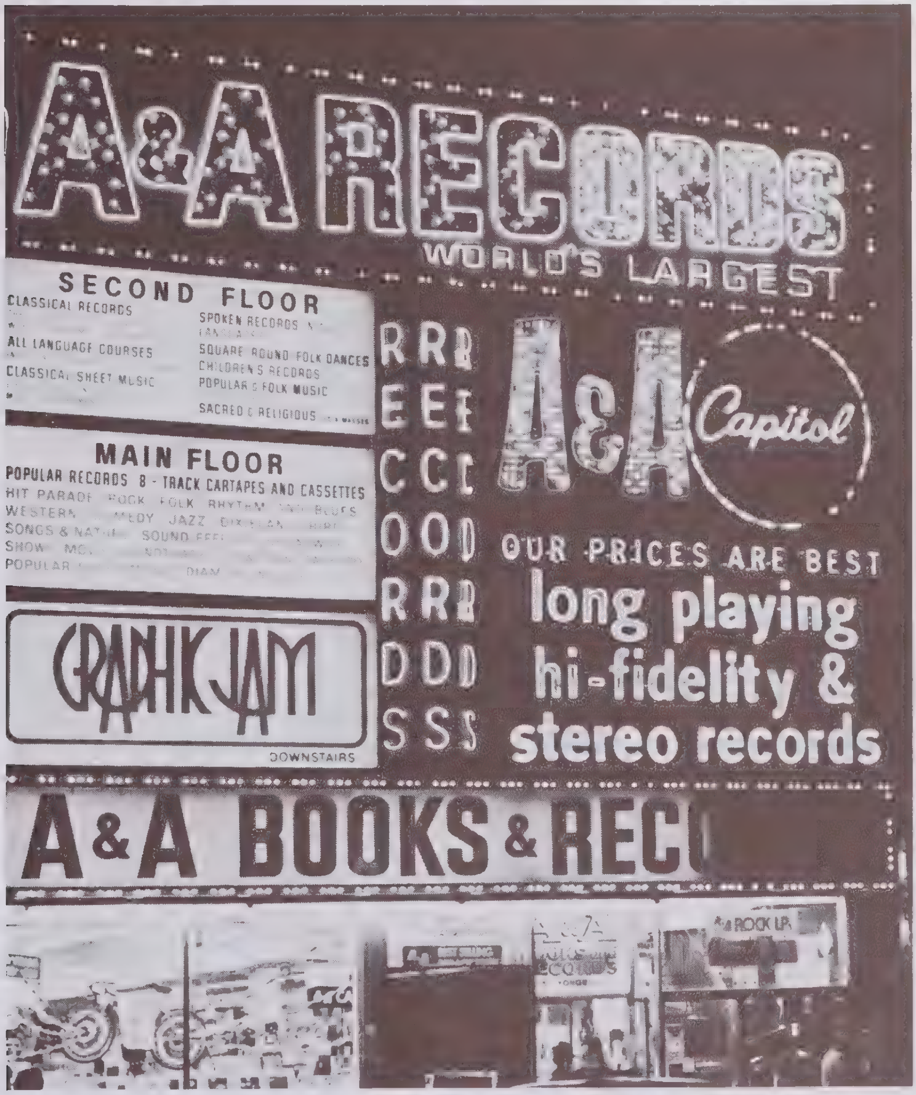
hand spectrometer
coloured pencils
gas discharge tubes

1. At several stations around the room you will find a gas discharge tube. These tubes are connected to a high voltage source.

Warning

Under no circumstances should you come closer than 20 cm to the tubes or the high voltage source.

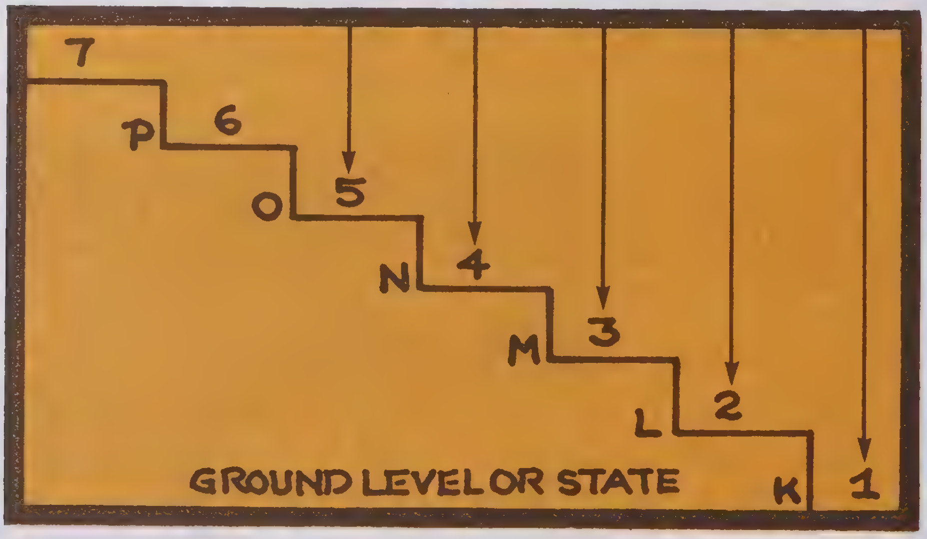
2. At each station record the name of the gas inside the tube.
3. Hold the spectrometer so that the plastic screen is toward you. Holding the screen very close to your eye, point the slit in the opposite end directly at the discharge tube. You should be able to see light from the discharge tube through the slit.
4. Here is the hard part. Holding the spectrometer very firmly let your eye turn about 30° either way. The pattern you see is called a *spectrum*. Try to draw this spectrum. Two things are important in your drawing:
   (a) the colour;
   (b) the spaces between the bands.
5. Repeat this procedure for any other tubes that are available.


A&A RECORDS
WORLD'S LARGEST
SECOND FLOOR
CLASSICAL RECORDS
ALL LANGUAGE COURSES
CLASSICAL SHEET MUSIC
SPOKEN RECORDS
SQUARE ROUND FOLK DANCES
CHILDREN'S RECORDS
POPULAR & FOLK MUSIC
SACRED & RELIGIOUS
MAIN FLOOR
POPULAR RECORDS 8 - TRACK CARTAPES AND CASSETTES
GRAPHIKJAM
DOWNSTAIRS
RECORDS
A&A
Capitol
OUR PRICES ARE BEST
long playing
hi-fidelity &
stereo records
A&A BOOKS & REC

Questions

1. How do the spectra of different substances compare?
2. What does this tell you about the energy staircase in different atoms?
3. Are the electrons that you see going up the stairs or down the stairs? Explain.
4. What is the function of the high voltage source?

About the only difference in our analogy is that the electrons are first pushed to the top of the stairs by the electrical (or some other form of) energy. We see them falling down the stairs with the spectrometer. Each band of light represents a different step. We might try to draw a diagram of it this way.



3.8

Biological Implications (Optional)

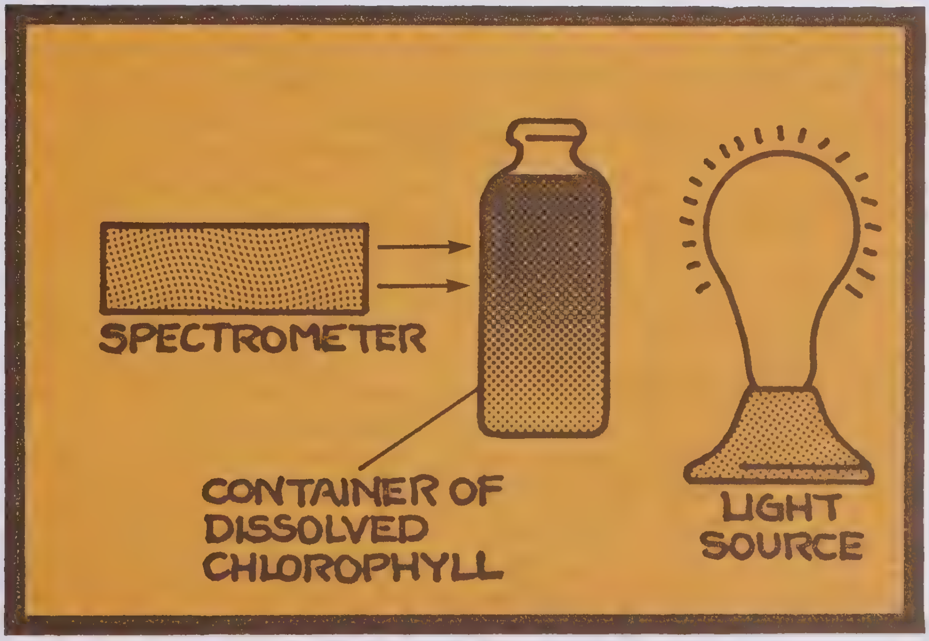
Some of you may wish to do this investigation as a project or in conjunction with your biology program. Just as physics is frequently basic to chemistry, so chemistry frequently pops up in biology. One thing we must learn is that no science stands by itself.

hand spectrometer
green grass
methanol
flat-sided container, e.g. glass milk bottle, medicine bottle, Tic-Tac box
white light source
hot plate

1. Set up the spectrometer and the light source as shown in the diagram.
2. View and record the spectrum of the white light produced by the source.
3. Very carefully warm several handfuls of fresh green grass in a flask with 200 ml of methanol. Continue the heating until the liquid is a deep rich green colour.
4. Cool the liquid. Pour off some of the liquid into the flat-sided container. This is your sample cell.
5. Place the sample cell between the light source and the spectrometer. View the transmitted light and record the spectrum produced.
6. Compare the spectrum produced by the light source to the spectrum produced by the chlorophyll. The spectrum produced by a discharge tube is called an *emission spectrum*. The chlorophyll spectrum is called an *absorption spectrum*.

Questions

1. What explanation do you have for the difference between the chlorophyll and the white light spectra?
2. What implications does your explanation have as far as photosynthesis is concerned?
3. Did you know there are two kinds of chlorophyll? What implications does this have for your explanations?
4. In the Independent Study Projects you will find a topic called chromotography. Perhaps you could combine chromotography with absorption spectra.



Law of Atomic Number

Bernard Joffe, in *Moseley and the Numbering of the Elements*, writes:

> It has been the fate of some men to accomplish in their young manhood, a work of surpassing importance and then to have a brilliant career cut short by a great and perhaps needless catastrophe. Such was the fate of Henry Moseley, whose life work encompassed less than four years. The world had hardly heard of him before he was gone.

Moseley had joined Ernest Rutherford at Manchester University in 1910. He worked on his research for four short years, then joined the British Army as a volunteer at the outbreak of World War I in 1914. In August of 1915 he was killed by a Turkish bullet through the head, the victim, along with thousands of others, of one of the greatest military blunders of all times. Much has been written about the brilliance of Moseley's work but perhaps Robert Millikan wrote the best epitaph: "Had the European War had no other result than the snuffing out of this young life, that alone would make it one of the most hideous and irreparable crimes in history."

It was Moseley who determined that as each radium atom transmutated, a single electron was ejected. It was Moseley who measured the half-life of actinium and also measured the charge buildup on the radium residue as it lost electrons.

But in his most important work he used the legacies of several men to break new ground. Wilhelm Roentgen had discovered x-rays. Max von Laue had discovered that crystals cause x-rays to split into distinct patterns. William Henry Bragg and his son, William Lawrence Bragg, had shown what von Laue's patterns had meant in terms of the crystal structure of the target. Ernest Rutherford, who gave us the nuclear atom, theorized that the charge on the nucleus should be proportional to the atomic mass.

Moseley, working with incredible speed for a scientist, absorbed all these ideas, digested them and devised a new experiment. He produced x-rays from a metal target and then measured the frequencies of the x-rays produced. He then changed the metal in the target again and again until he had done this for about 40 elements. He discovered that the heavier elements produced higher frequencies. The greater the atomic mass, the higher the frequency of x-rays produced.

His conclusion was simple: "There is in the atom a fundamental quantity which increases by regular steps as we pass from each element to the next. This quantity can only be the charge on the central positive nucleus."

This was Moseley's Law of Atomic Numbers. It was an idea so simple and so basic that it provided a known direction into the unknown world of undiscovered elements.

This is what it means. The atom with the lowest atomic mass, hydrogen, has the lowest nuclear charge. That would be one unit of charge or one proton. Next would be helium with two units or two protons. Next lithium with three protons.

The number of protons in the nucleus is called *atomic number*. It must be a whole number. It is symbolized by Z.

Table 3.3

	Z *Atomic Number*	*Atomic Mass*
Hydrogen	1	1.0
Helium	2	4.0
Lithium	3	6.9
Beryllium	4	9.0
Boron	5	10.8
Carbon	6	12.0
Nitrogen	7	14.0
Oxygen	8	16.0
Fluorine	9	19.0
Neon	10	20.2

Consider what we now know about the atom.

1. It is neutral.
2. The positive charge (protons) is located in the nucleus.
3. The electrons are in orbit around the nucleus.
4. The number of protons in the nucleus is equal to the atomic number of the element. All atoms of the same element have the same atomic number. All atoms of the same atomic number are the same element.

Hydrogen, atomic number 1, has one proton. Since it is neutral it must also have one electron. Remembering that a proton has a mass of 1 u and the electron has a mass of 1/2000 u, the total mass of the hydrogen atom will then be:

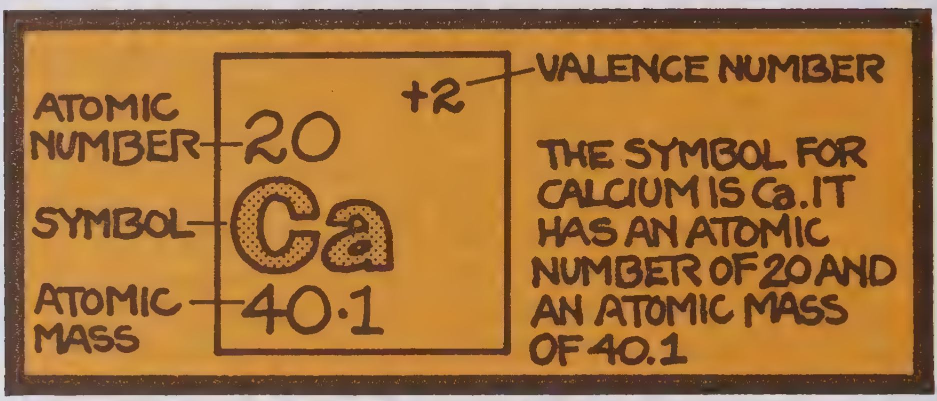
$$1 \text{ u} + \frac{1}{2000} \text{ u} = 1.0005 \text{ u}$$

This corresponds favourably with the figure of 1.0 which appears in Table 3.3.

You can find the atomic number of any element by looking at the periodic table inside the cover of this text. It is the number above and to the left of the symbol of the element.

The atomic mass is found in the same table and is the large number directly below the symbol for the element. (Decimals may be rounded off for this course.)

With the table then we can find the symbol, the atomic number and the atomic mass. Consider calcium as an example:


ATOMIC NUMBER
20
+2
VALENCE NUMBER
SYMBOL
Ca
ATOMIC MASS
40·1
THE SYMBOL FOR CALCIUM IS Ca. IT HAS AN ATOMIC NUMBER OF 20 AND AN ATOMIC MASS OF 40.1

Let us now consider the second element, helium. The atomic number is 2, therefore helium has two protons which have a mass of 2 u. Helium is neutral and therefore has 2 electrons which have a mass of 2/2000 u. The total is therefore:

$$2 + \frac{2}{2000} = 2.001 \text{ u}$$

But the mass listed in the table is 4.0 u. Something is wrong. Let's check on lithium. The atomic number is 3.

$$\begin{aligned} 3 \text{ protons} &= 3 \text{ u} \\ 3 \text{ electrons} &= \frac{3}{2000} \text{ u} \\ \text{total} &= 3.0015 \text{ u} \end{aligned}$$

The atomic mass is listed as 6.9 u.

Try to suggest a possible explanation for these discrepancies in our calculations.

You are absolutely right!

There is something in the atom that we have missed. We even know two things about it, maybe even three.

1. It has mass.
2. It must, therefore, be in the nucleus.
3. It must be neutral or have no charge at all.

Henry Moseley made these deductions, but it was not until 1932 that Chadwick, his friend and co-worker from Rutherford's laboratory in Manchester, actually discovered these new sub-atomic particles. They were given the name *neutrons*.

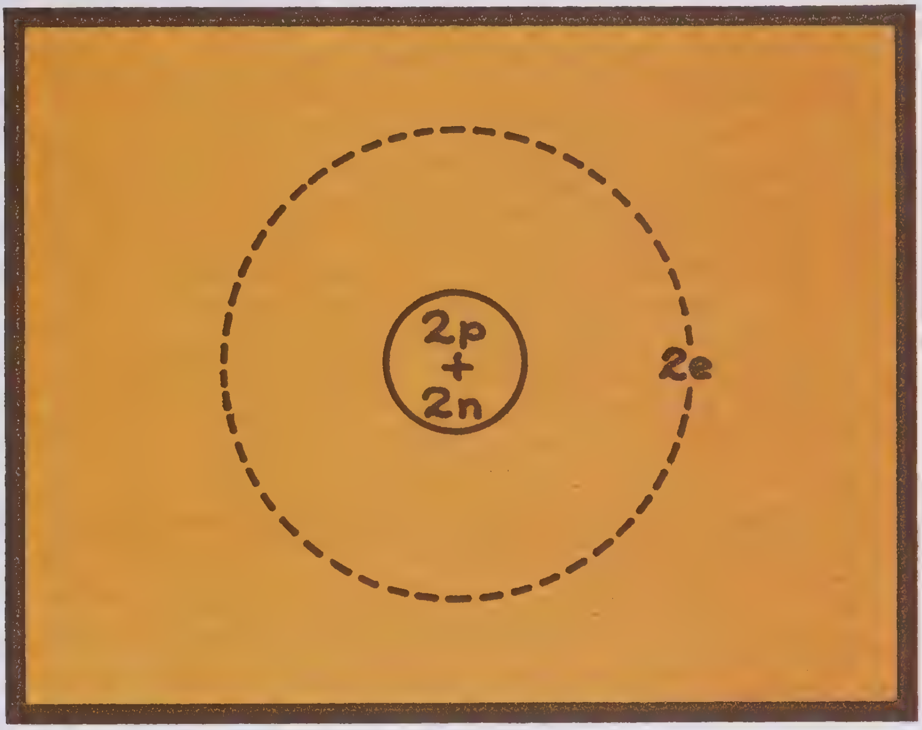
The neutron was found to have the same mass as a proton, but it carried no charge.

Table 3.4

Particle	*Mass*	*Charge*	*Symbol*
hydrogen atom	1 u	none	H
electron	1/2000 u	negative (−1)	e−
proton	1 u	positive (+1)	p+
neutron	1 u	none	n

If we now go back to our helium atom with its atomic number of 2, we see that the sum of protons and electrons are 2.001, let's say 2.0, while the atomic mass is 4.0 u. We, therefore, have 2.0 u of unaccounted for mass. This must represent the neutrons. One neutron has a mass of 1 u, therefore 2 u represents two neutrons.

Our helium atom might be represented this way.



For lithium, the atomic number is 3. Therefore lithium has 3 protons. Since electrons have so little mass, we can ignore their mass and say that the mass of protons and electrons is 3.0 u. The listed atomic mass is 6.9. We will round it off to 7 for the moment.

AM − Ƶ = mass of neutrons
7.0 − 3.0 = 4.0 u

One neutron has a mass of 1 u, therefore 4 u represents four neutrons. So lithium nucleus contains three protons and four neutrons.

Let's do calcium.

Calcium atoms have 20 protons and to balance that charge, 20 electrons. These would add up to 20 + 20/2000 u, that is, about 20 u. The total mass is 40 u.

AM − Ƶ = mass of neutrons
40 − 20 = 20 u

One neutron has a mass of 1 u so 20 neutrons have a mass of 20 u. Calcium has 20 protons, 20 electrons, and 20 neutrons.

In general we find the number of neutrons by subtracting the atomic number from the atomic mass.

Neutrons = atomic mass − atomic number.

Copy Table 3.5 into your notebook and complete it.

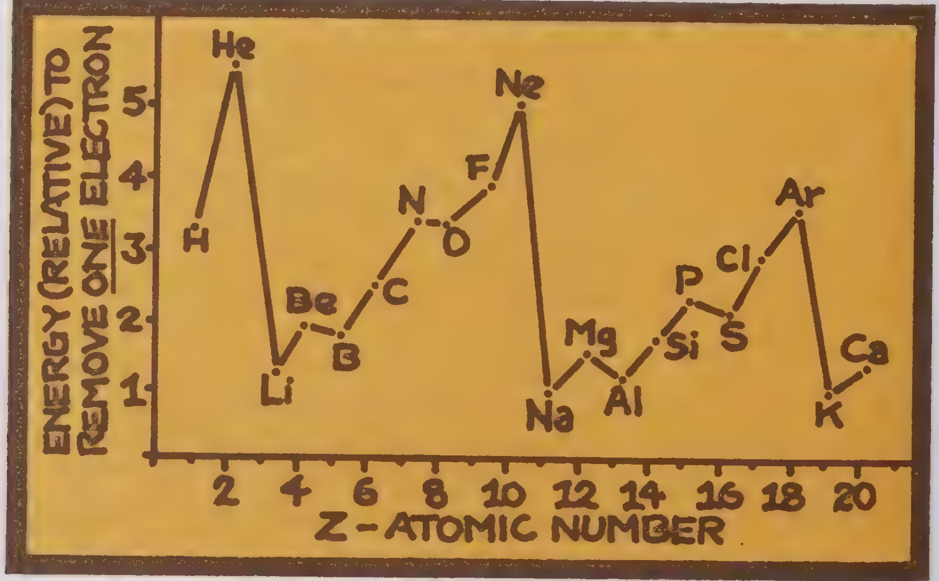
Table 3.5

	Z	No. of Protons	No. of Electrons	AM	No. of Neutrons
Hydrogen	1	1	1	1	
Helium	2	2	2	4	
Lithium	3	3	3	7	
Beryllium					
Boron					
Carbon					
Nitrogen					
Oxygen					
Fluorine					
Neon					
Calcium	20	20	20	40	20
Uranium	92				
Nobelium	102				

Electron Configurations

We have already seen that by using energy it is possible to remove electrons from an atom. That was what happened when you walked across the wool rug and got an electric shock when you reached for the doorknob. Scientists in various places have measured the amount of energy required to remove first one electron from an atom, then a second electron, then a third, etc.

When the energy to remove a single electron from each of the first 20 elements is plotted on a graph it produces this curve:



Consider hydrogen ($Z = 1$) with its proton attracting one electron. The energy needed to remove that electron will be equal to the work done in pulling it away from the nucleus. Very roughly, then, the energy required will be proportional to the pull required. For hydrogen we must pull with about 3 units of pull to overcome the attraction between the nucleus and the electron.

Consider helium ($Z = 2$) with its two protons tugging on each electron. The pull is about twice as great, roughly 5.5 units. It is not quite twice because helium's second electron is helping to push the first one out, or so it seems.

If we attempt to make a prediction for lithium we would expect that three protons' pull would be about 9 units of pull minus the repulsion of two electrons — maybe 1 or 2 units. That would give about 7 units. Measurement of the actual pull gives about 1 unit.

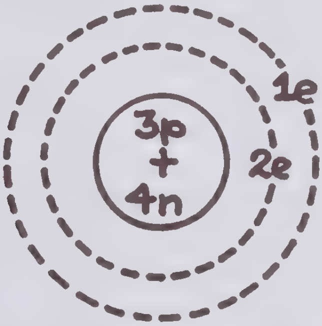
How can you explain that? We'll give you a hint.

"The electrostatic force of attraction (or repulsion) between two charges is directly proportional to the product of the two charges and inversely proportional to the distance between them squared."

$$F \alpha \frac{q_1 q_2}{d^2}$$

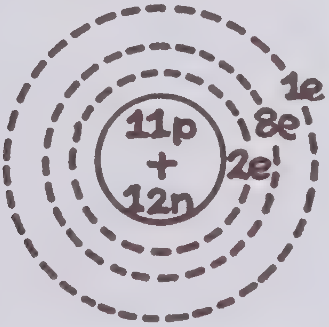
It would appear that lithium's third electron is in a different orbit from the first two. If it were moving on a bigger radius the force would be much smaller.

We might picture lithium this way:

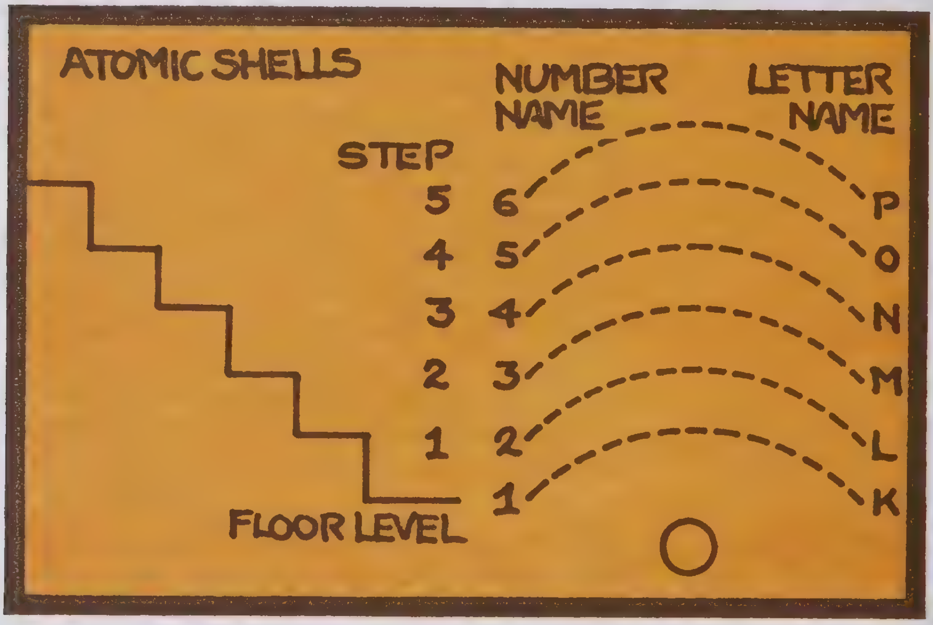


If we ignore the small bumps in the graph for the moment, it would seem as though this new orbit can hold the electrons for beryllium boron, carbon, nitrogen, oxygen, fluorine and neon.

At this point the second orbit is also "full," and a new orbit begins for sodium. We might represent it in this way:



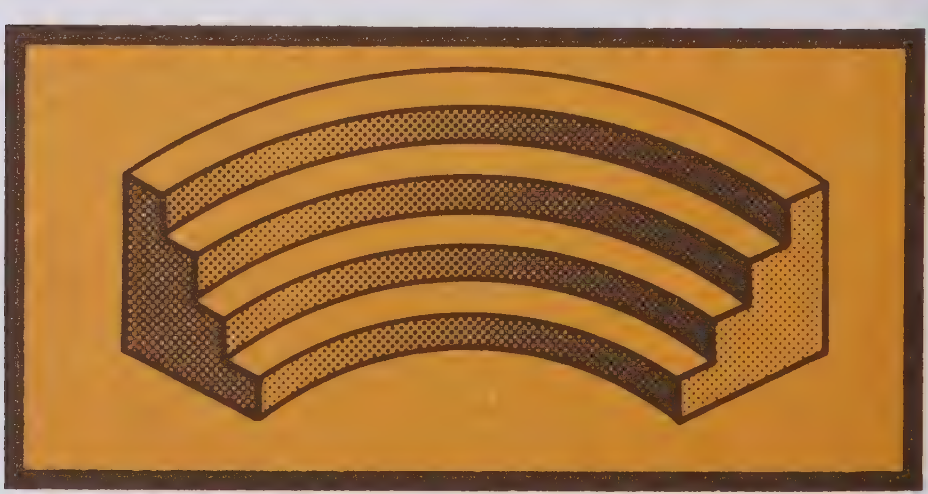
These new orbits were called *shells* or *energy levels*. How well they fit in with the ideas of Neils Bohr mentioned earlier. Bohr had used them to explain spectral lines. Now we see that energy levels are also useful in explaining electron removal energies.



The phrase "number name" is not a very good one. Scientists have decided to use the phrase "principal quantum number" instead. This means that an electron in the K shell has a principal quantum number of 1. An electron in the M shell would have a principal quantum number of 3.

It appears that there is a limit to the number of electrons each shell can hold. Wolfgang Pauli of Austria did much of his work in the area of electron populations.

Earlier we likened the shells to the steps of a staircase. If we asked 20 people to assemble on a staircase it would be very crowded if everyone stood on the bottom step. If you were a photographer how would you arrange a large group on these steps?



It should be apparent that the higher we go the more people we will place on each step. So it is with the atom. The first shell holds two electrons. The second holds eight. The third holds 18, then 32, then 50, and so on. These numbers are related in a very simple formula:

$$\text{Maximum occupancy} = 2n^2$$

where n is the shell number or principal quantum number.

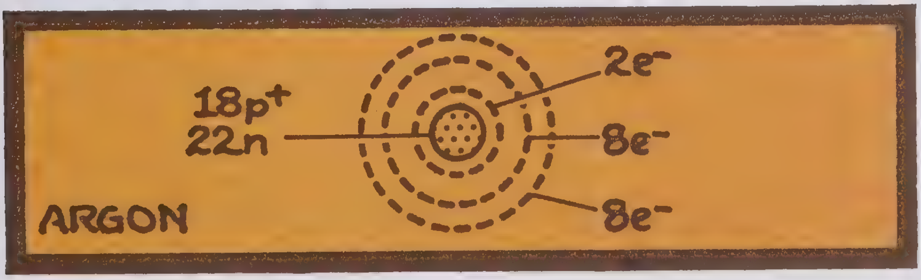
It is very important to note that this equation gives the *maximum* occupancy. A shell can be only half full, or even empty. As we have seen with hydrogen, there is only one electron so the first shell is only half full and all the rest are empty under normal circumstances. If hydrogen is excited, then that single electron may move into the second, third or fourth shell briefly before returning to the ground state.

Generally speaking the first shell fills before the second. The second shell fills before the third, and so on. We can now begin to build up a picture of electron location within the shells for electrons in the ground state.

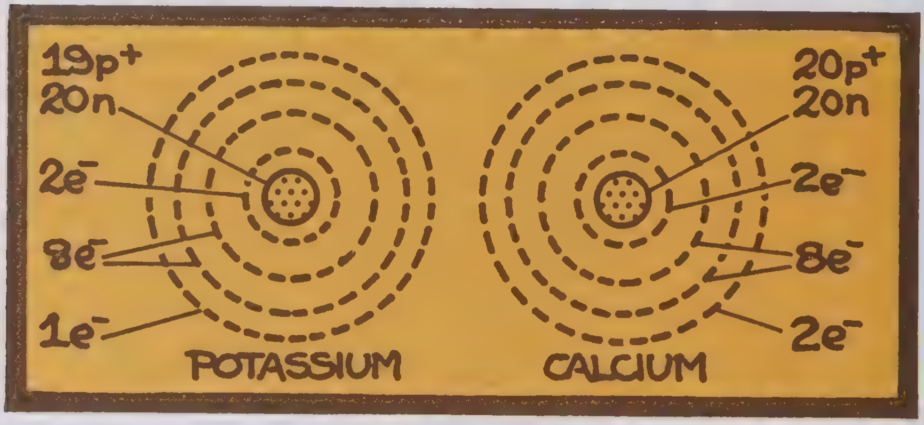
Table 3.6

Atomic Number	*Element*	*1st Shell Population*	*2nd Shell Population*	*3rd Shell Population*
1	hydrogen	1		
2	helium	2		
3	lithium	2	1	
4	beryllium	2	2	
5	boron	2	3	
6	carbon	2	4	
7	nitrogen	2	5	
8	oxygen	2	6	
9	fluorine	2	7	
10	neon	2	8	
11	sodium	2	8	1
12	magnesium	2	8	2
13	aluminum	2	8	3
14	silicon	2	8	4
15	phosphorus	2	8	5
16	sulphur	2	8	6
17	chlorine	2	8	7
18	argon	2	8	8

A simple atomic diagram for argon might look like this:



It wouldn't be right to let you think it was all this simple and neat. Consider potassium and calcium:



As you can see, the fourth shell begins to fill before the third shell is full. The regularity that we have been following ends with argon. For atoms beyond argon it is simpler to consult a table. Table 3.7 shows the structure of some of the more common elements beyond calcium.

Table 3.7

		Electron Occupancy in Shell Number					
Atomic Number	*Element*	*1*	*2*	*3*	*4*	*5*	*6*
22	titanium	2	8	10	2		
24	chromium	2	8	13	1		
25	manganese	2	8	13	2		
26	iron	2	8	14	2		
27	cobalt	2	8	15	2		
28	nickel	2	8	16	2		
29	copper	2	8	18	1		
30	zinc	2	8	18	2		
33	arsenic	2	8	18	5		
35	bromine	2	8	18	7		
36	krypton	2	8	18	8		
47	silver	2	8	18	18	1	
48	cadmium	2	8	18	18	2	
50	tin	2	8	18	18	4	
51	antimony	2	8	18	18	5	
53	iodine	2	8	18	18	7	
54	xenon	2	8	18	18	8	
56	barium	2	8	18	18	8	2
80	mercury	2	8	18	32	18	2
82	lead	2	8	18	32	18	4
86	radon	2	8	18	32	18	8

It is with some regret that we have to tell you there is no simple formula to produce these occupancy numbers. While it may look complex, it has been investigated thoroughly and there are some interesting ideas that we might add just to whet your appetite.

Pauli and others have shown that within each shell there are sub-shells. These sub-shells allow us to explain many new and exciting things in chemistry. For the moment we will stick with the model that we have been working on.

Given a periodic table you should now be able to determine the number of protons, neutrons and electrons in any atom and in addition be able to locate the electrons for the first 20 elements.

Problems

1. Name the three sub-atomic particles we have studied. Give the charge and mass for each particle.
2. Describe how to determine the number of protons, neutrons and electrons in any atom.
3. How is the maximum occupancy of a shell determined?
4. Draw a simple diagram to represent a single atom of:
   (a) titanium
   (b) copper
   (c) tin
5. What conditions must be fulfilled for an atom to be:
   (a) electrically neutral;
   (b) carry a negative charge;
   (c) carry a positive charge?
6. (a) Give the electron arrangements for each of the inert gases.
   (b) What do these electron arrangements all have in common?
   (c) Why do you suppose these elements are called *inert*?

Every element is different from all other elements in more or less distinctive ways. No one could mistake sulphur for copper or carbon for aluminum. While silver and platinum may look alike, there are many physical and chemical properties that enable us to tell them apart. So it is with all the elements.

We have just proposed that the basic difference in the atoms of different elements lies in the number of sub-atomic particles that make them up. If this is true, then knowing the number and arrangement of the protons, neutrons and electrons should tell us something about the atom.

Question 6 above might have been answered in the form of a table like Table 3.8.

Table 3.8

	Shell Number					
Element	*1*	*2*	*3*	*4*	*5*	*6*
helium	2					
neon	2	8				
argon	2	8	8			
krypton	2	8	18	8		
xenon	2	8	18	18	8	
radon	2	8	18	32	18	8

These inert gases all have several things in common:

1. All are gases at room temperature.
2. All are composed of monatomic molecules.
3. All liquify at very low temperatures.
4. All are very unreactive (inert) under ordinary conditions.

We might ask why should this be so? Why do some atoms combine to form molecules whereas the inert gases do not?

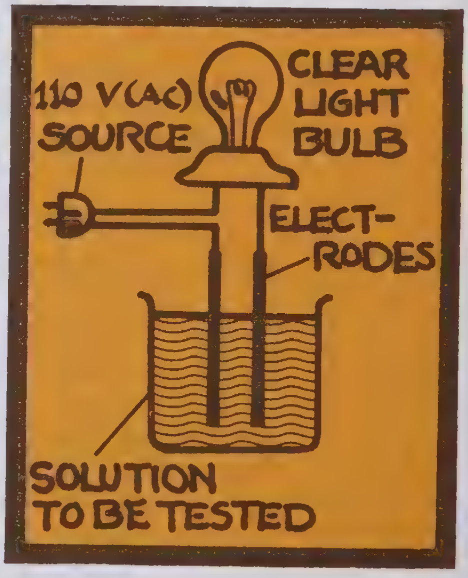
Hydrogen atoms combine to form diatomic hydrogen molecules. Can we conclude that the two hydrogens are attracted to each other? Attracted! That word has come up before in this section. Could it have something to do with the protons and electrons in each atom? We did miss one other thing the inert gases have in common. They all have eight electrons in the outer occupied energy level. (Except for helium, but that first shell is full with two.) Could it be that these eight electrons make these gases unreactive? We need more information to work with.

Conductivity of Solutions

3.9 Conductivity

Each group will be required to prepare 100 ml of a certain solution or substance in a 250 ml beaker. Each group will be assigned a different substance such as:

(a) pure water (distilled)
(b) tap water
(c) methanol
(d) glacial acetic acid
(e) 1:1 acetic acid in water
(f) 1:10 acetic acid in water
(g) sodium iodide in water
(h) potassium bromide in water
(i) calcium chloride in water
(j) sugar in water
(k) any other substance as directed You will also need a conductivity tester.



Warning

Do not touch the electrodes once the conductivity tester has been plugged in.

1. Prepare your substance in a carefully labelled beaker.
2. Set up the conductivity tester as shown in the diagram. When you have tested your own substance leave the apparatus assembled and move around the room to other stations and record the results of other groups.

Questions

1. What conditions must be satisfied before the bulb can light up?
2. Why do you think that some substances such as silver, copper, or aluminum can conduct electricity?
3. Why do you think that pure water will not conduct (much) electricity?
4. What do you think happens when a substance like sodium iodide is dissolved in water that allows the current to pass through the solution?
5. Can you suggest an explanation for the conduction of electricity by solutions such as acetic acid or calcium chloride in water but not by sugar in water or pure methanol?

It would seem further investigation is required.

Let's consider the results of this experiment just for a moment. You may have found that some of these solutions conducted and some did not. Let's ignore those that did not for a minute and consider those that did.

You should remember from your physics that an electrical current is a flow of electrons. That means that, in the conducting solutions, electrons must be entering the solution at one side and leaving the solution at the other. What is happening in between the electrodes?

Here are four possible explanations. Try to add at least two more.

1. Electrons swim from one electrode to the other.
2. Electrons are passed from one molecule to another across the solution.
3. Molecules line up and form a bridge for the electrons to pass over.
4. Electrons attach themselves to a molecule and are carried across by that molecule.
5.
6.

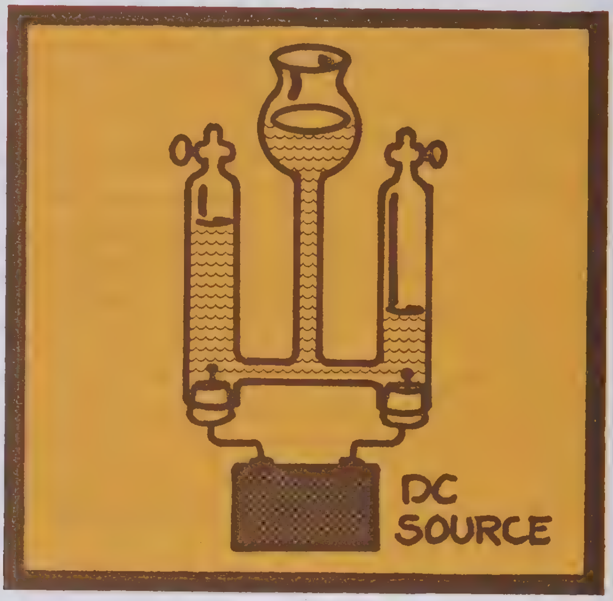
What happens to these explanations when we remember that some of the solutions do not conduct electricity?

3.10

Electrolysis

Hoffman apparatus
test tubes
cedar splint

hydrochloric acid



1. This experiment requires us to pass *direct* current through a solution that conducts electricity. An apparatus that can be used for this type of experiment is the Hoffman apparatus. Any conducting solution may be used. For this experiment we will use hydrogen chloride dissolved in water (hydrochloric acid).
2. After the appartus has been running for some time note carefully any changes in the colour and level of the liquid in each arm of the apparatus.
3. Try to identify the colourless gas by collecting some in an inverted test tube and testing with a burning splint.
4. Try to identify the coloured gas by:
   (a) smell;
   (b) its effect on moist blue litmus.
5. Be sure you record the polarity of the two electrodes.

Questions

1. What gas was collected over the negative terminal of the apparatus?
2. What gas was collected over the positive terminal of the apparatus?
3. What can you say about the charge on the hydrogen in hydrogen chloride if it moves to the negative terminal?
4. What must happen to the hydrogen in hydrogen chloride at the negative terminal to turn it into neutral hydrogen atoms?
5. What can you say about the charge on the chlorine in hydrogen chloride if it moves to the positive terminal?
6. What must happen to the chlorine in hydrogen chloride at the positive terminal to turn it into neutral chlorine atoms?

Aluminum reduction plant at Kitimat, BC. Aluminum oxide is reduced electrolytically to metallic aluminum in plants such as this. *Courtesy of Aluminum Company of Canada, Ltd.*

7. What must happen to a hydrogen chloride molecule when it is put into water to form hydrochloric acid?

Now consider the possible explanations for conductance in the light of this new information. Which of our explanations seems the most reasonable now?

Ions

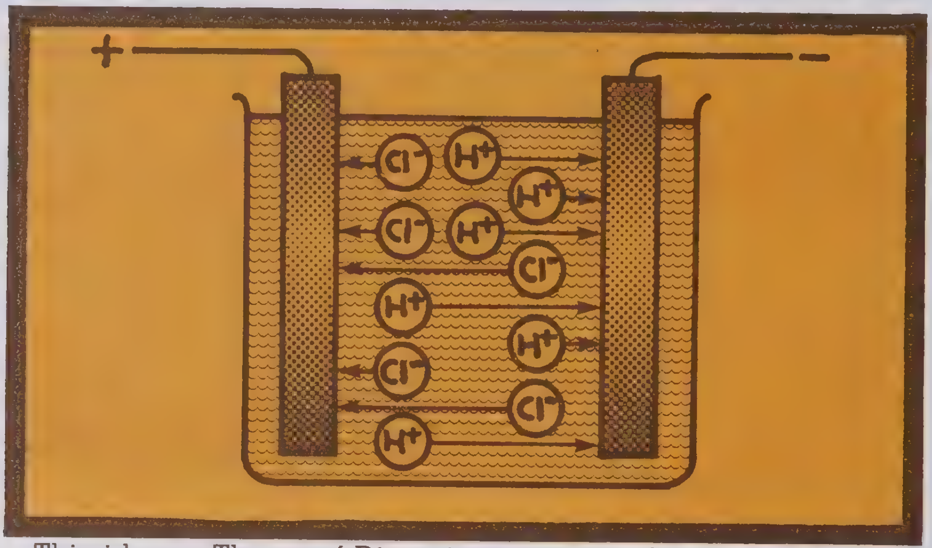
Any attempt to explain Experiments 3.9 and 3.10 must include explanations for both the conductors and the non-conductors. One such explanation was proposed some time ago by a Swedish chemist named Svante Arrhenius in about 1885, although it was neither accepted nor recognized until much later.

Arrhenius suggested that under certain conditions some molecules could split or dissociate into two or more charged pieces which he called *ions* (from a Greek word which means "wanderer"). These charged ions were then free to move — if they were in the liquid or gaseous state — in the same random way a molecule did. If, however, electrodes were placed near the ions, the positive ions would move toward the negative electrode and the negative ions would move toward the positive electrode. In our experiments this means that hydrogen ions must be positive and chloride ions must be negative. We will symbolize a hydrogen ion as H^+ and a chloride ion as Cl^-.

When the hydrogen ion reaches the negative electrode it accepts an electron from it and becomes a neutral hydrogen atom:

$H^+ + e^- \longrightarrow H^0$

Similarly, when the chloride ion reaches the positive terminal it gives up its extra electron and becomes a neutral chlorine atom: $Cl^- - e \rightarrow Cl^0$

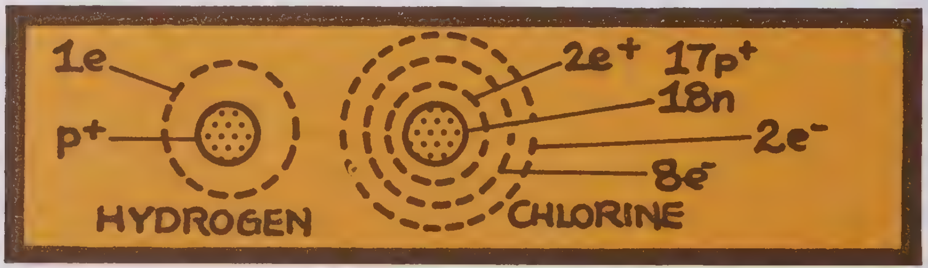

+
−
Cl⁻
H⁺
H⁺
Cl⁻
H⁺
Cl⁻
H⁺
H⁺
Cl⁻
Cl⁻
H⁺

This idea or Theory of Dissociation can explain quite nicely:

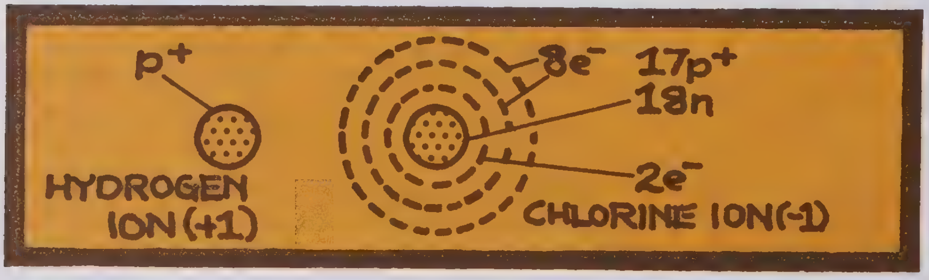
1. why some solutions conduct electricity;
2. why some solutions do not conduct electricity;
3. the products of electrolysis

with the simple premise that only some molecules dissociate into charged particles called ions.

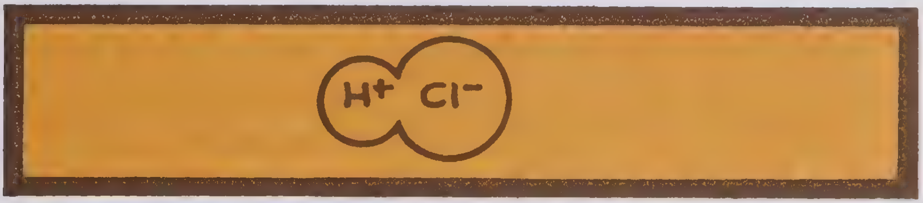
But how do neutral atoms become ions in the first place? We must look at hydrogen chloride again to answer this. The structure of the two atoms is:


1e
p+
HYDROGEN
2e+
17p+
18n
2e-
8e-
CHLORINE

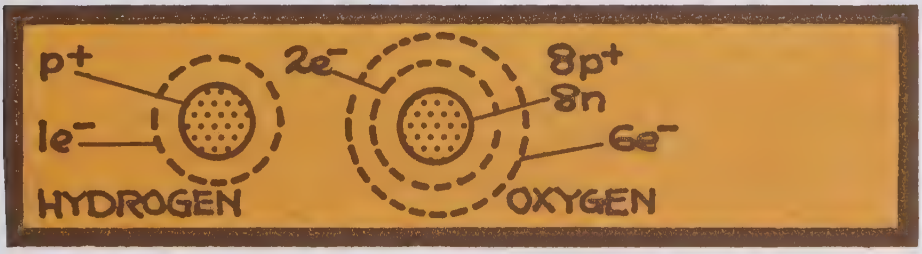
Neutral hydrogen can become positive by giving up its one electron. Neutral chlorine can become negative by gaining one more electron. If hydrogen gives its electron to chlorine, then both neutral atoms become charged ions and the resulting ions are:


p+
HYDROGEN ION (+1)
8e-
17p+
18n
2e-
CHLORINE ION(-1)

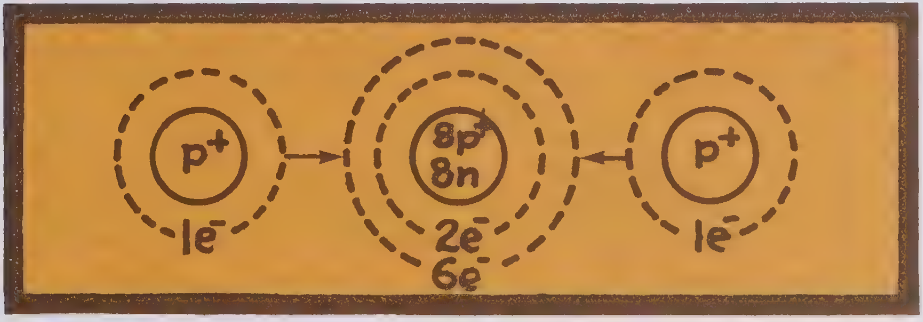
Since they have opposite charges they will attract each other and form a molecule of hydrogen chloride which might look like this:



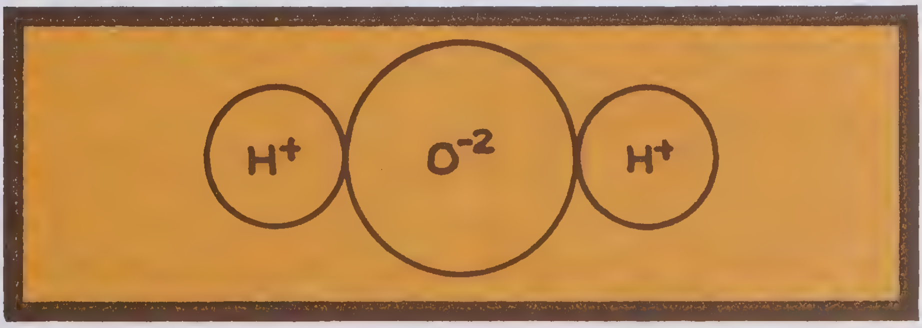
Consider the two substances, hydrogen and oxygen:



Each hydrogen proton attracts its own electron. It also attracts all other nearby electrons although not as vigorously. Oxygen's eight protons attract its own eight electrons and any others that happen to be nearby. If oxygen is mixed with hydrogen then oxygen will attract hydrogen's electrons. If oxygen can attract them more than hydrogen can, then oxygen will be able to take them from hydrogen. Since oxygen has six electrons in its second shell and since the shell will hold eight electrons, there is space available for two more electrons. Two electrons may be captured, one from each of two hydrogen atoms.



This will leave the hydrogen as a single proton with a positive charge. It will give oxygen ten electrons and eight protons or a net charge of two negative. Positive attracts negative. The hydrogen is held to the oxygen in a 2:1 ratio. Remember? H_2O.



The particles that form the water molecule are atoms that have gained or lost electrons. Such charged particles are called *ions*. Atoms that give up electrons will become positive ions. Atoms that take on electrons will become negative ions.

Remember the inert gases? They each had eight electrons in the outer shell (except for helium which had a complete first shell). Now we see that chlorine took electrons from hydrogen to get eight electrons in its outer shell. So did oxygen.

Chemists have discovered that eight electrons in the outer shell seems to be a desirable state of affairs for an atom. They have also determined that a full outer shell is desirable state for an atom. Put the two ideas together and a valuable rule of thumb emerges: Atoms will gain, lose or share electrons so as to achieve an electron population of either eight or $2n^2$ electrons in the outer shell. This condition seems to make an atom stable. Atoms that have achieved this condition seem content with their lot and do not actively seek to change it. (There are always a few exceptions but we will ignore them for now.)

In simple terms it means that atoms that have nearly eight electrons or nearly a full shell will attract electrons from wherever they can so as to get eight electrons or a full shell. If you go down Table 3.9 you will find that

Table 3.9

Element	*Has (in outer shell)*	*Needs*	*To Make*
fluorine	7 electrons	1 electron(s)	8 electrons
chlorine	7	1	8
sulphur	6	2	8
oxygen	6	2	8
phosphorus	5	3	8
nitrogen	5	3	8
carbon	4	4	8
silicon	4	4	8

If you look at carbon and silicon carefully you will see that they could also reach stability by losing four electrons. Nitrogen and phosphorus could reach stability by losing five electrons. Similarly, sodium could lose one, magnesium two, aluminum three and fall back, as it were, to eight electrons in the outermost occupied shell.

Valence Numbers

The number of electrons that an atom could gain, lose or share to achieve stability is called its combining capacity and is represented by a number called the *valence number*.

Sodium may give up one electron and become a positive ion. It therefore has a valence number of +1.

Magnesium may give up two electrons and become a positive ion. It therefore has a valence number of +2.

Chlorine may accept one electron and become a negative ion so it has a valence number of −1.

Sulphur may accept two electrons and become a negative ion so it has a valence number of −2.

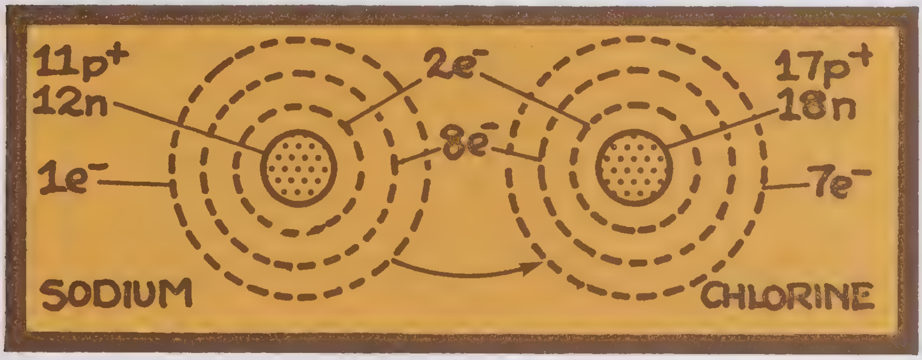
The size of the valence number tells you the number of electrons involved; the sign tells you whether they were lost (+) or gained (−).

Consider the following example:

Sodium has a very weak grip on its outer electron — it is easily removed to make sodium a +1 ion (written Na^+).

Chlorine has space for one more electron and exerts a strong attraction on sodium's electron; if the electron is transferred chlorine will become a chloride ion with a −1 charge (Cl^-).

The exchange would look like this:



The resulting ions would be held together by their unlike charges to form a molecule of sodium chloride.

The force of attraction that holds two atoms together is called a *bond*. The force that holds the sodium chloride molecule together is between two charged ions and so is called an *ionic bond*.

It follows then since ions are formed when one atom takes electrons from another, substances formed this way would be called *ionic substances*.

In symbolic form we represent the reaction this way:

$Na^0 \longrightarrow Na^+ + e^-$

$Cl^0 + e \longrightarrow Cl^-$

$Na^0 + Cl^0 \longrightarrow Na^+ + Cl^-$

The Na^+ and Cl^- are held together by an ionic bond to form NaCl — common table salt.

$Na^0 + Cl^0 \longrightarrow NaCl$. This symbolic representation of a chemical reaction is called an equation. The hydrogen chloride we saw earlier would form this way: $H^0 + Cl^0 \longrightarrow H^+ + Cl^-$ but H^+ and Cl^- would be held together by an ionic bond to form HCl — hydrogen chloride. Both of these ionic substances can dissociate when dissolved in water to allow the ions to wander.

You can find the valence number of an element in one of three ways:

1. Remember it.
2. Look it up on the periodic table.
3. Figure it out from the electron arrangement.

If we know the valence number of an element we can determine a possible formula for molecules that it may form with other elements.

Example 1

sodium — Na — valence +1

Sodium will tend to combine only with elements that can accept an electron. Those would be elements with a negative valence such as fluorine.

fluorine — F — valence −1

We always write the positive ion first followed by the negative ion:

$Na^{+1}\ F^{-1}$

The resulting molecule must be neutral which means that the sum of the charges on the ions must be zero:

$Na^{+1}\ F^{-1}$ — $+1-1 = 0$

Therefore the molecule contains one sodium ion and one chloride ion which we would write like this:

$Na_1\ F_1$

However when we write formulae the subscript $_1$ is not used. The absence of a subscript is understood to represent $_1$. The final form of the formula is:

NaF — sodium fluoride

Notice that the valence numbers do not appear in the formula.

Without trying to confuse you we must make one other point.

Na_2F_2

would also represent a neutral molecule. Generally we always use the formula with the smallest subscripts. We must hasten to add that there are exceptions to this that you will see later. For now, always reduce your formula to the smallest set of subscripts possible.

Example 2

potassium — K^{+1}

oxygen — O^{-2}

First the positive, then the negative:

$K^{+1}\ O^{-2}$

So that the sum of the charges will be zero we will need two potassium:

$K_2^{+1}\ O^{-2}$ — $+1+1-2 = 0$

This gives us the formula:

K_2O — potassium oxide

Example 3

sulphur — S^{-2}
aluminum — Al^{+3}

First the positive, then the negative:

$Al^{+3}S^{-2}$

To balance the charge we will require:

$Al_2^{+3}S_3^{-2}$ — $+3+3-2-2-2 = 0$

This gives:

Al_2S_3 — aluminum sulphide

Now write the formula for:

(a) sodium chloride
(b) potassium sulphide
(c) calcium oxide
(d) aluminum phosphide
(e) magnesium bromide
(f) hydrogen flouride
(g) lithium nitride
(h) iron oxide
(i) copper chloride
(j) nickel sulphide

You can see the problem that presents itself in this type of exercise. Some substances have two different valence numbers listed. How do you know which one to use? In order to make that point clear a Roman numeral is added to the name to indicate which valence number is to be used.

For example, if we wish to use iron with a valence of +3 we write:

Iron (III)

If it was iron with a valence of +2 we wanted to use, then it would be:

Iron (II)

When you see the name manganese (IV) oxide you know that we are using the manganese with its +4 valence number.

Try to write the following formulae:

(a) iron (II) oxide
(b) copper (I) sulphide
(c) nickel (III) bromide
(d) phosphorus (III) oxide
(e) phosphorus (IV) oxide
(f) cobalt (II) chloride
(g) chromium (III) fluoride

As far as we are concerned in this section of the course, no substance has more than one negative valence number. This means that:

Sulphur in sulphides is always −2
Chlorine in chlorides is always −1
Bromine in bromides is always −1
Iodine in iodides is always −1
Nitrogen in nitrides is always −3
Phosphorus in phosphides is always −3

Binary Compounds

We do not need the Roman numeral notation for substances with a negative valence number, at least not yet. Only positive ions with more than one valence number need to be followed with the appropriate Roman numeral.

As we saw in Unit 2, substances formed from only two elements are called *binary compounds*. A binary compound may contain two atoms as in

$NaCl$

It may contain 26 atoms as in

C_8H_{18}

There is no theoretical limit to the number of atoms in a binary compound. The limit is in the number of kinds of atoms.

Equations

The formation of a binary compound could be represented this way:

sodium + chlorine ⟶ sodium chloride

This simple statement represents the chemical reaction that took place between sodium and chlorine to produce the substance sodium chloride. The arrow is often translated to mean "reacts to produce." This kind of statement in chemistry is called a *word equation*.

If we translate the words into the symbolic language of chemistry the word equation becomes:

$Na + Cl \longrightarrow NaCl$

We learned some time ago that chlorine is diatomic so our equation must show this:

$Na + Cl_2 \longrightarrow NaCl$

Now we have another problem. The last equation has two chlorine atoms on the left side of the arrow but only one on the right side. This violates the Law of Conservation of Mass. Since each sodium chloride molecule has only one chlorine atom in it we must have two sodium chloride molecules to have two chlorine atoms.

$Na + Cl_2 \longrightarrow 2NaCl$

Now we have two sodium atoms on the right and only one on the left. Let's fix that up.

$2Na + Cl_2 \longrightarrow 2NaCl$

This equation obeys the Law of Conservation of Mass and is called a *balanced formula equation*. It is a simple, one-line summary of a chemical reaction.

Example 1

Following the same steps, let's do the combination of hydrogen and oxygen to form water.

1. Hydrogen and oxygen $\longrightarrow$ water
2. $H + O \longrightarrow H_2O$
3. $H_2 + O_2 \longrightarrow H_2O$
4. $H_2 + O_2 \longrightarrow 2H_2O$
5. $2H_2 + O_2 \longrightarrow 2H_2O$

Example 2

Hydrogen and chlorine combine to form hydrogen chloride.

1. $H + Cl \longrightarrow HCl$
2. $H_2 + Cl_2 \longrightarrow HCl$
3. $H_2 + Cl_2 \longrightarrow 2HCl$

Problems

Write a balanced formula equation for each of the reactions listed below.

(a) magnesium combining with oxygen.
(b) copper (I) combining with sulphur.
(c) copper (II) combining with sulphur.
(d) aluminum combining with bromine.
(e) mercury (II) oxide decomposing to form mercury and oxygen gas.
(f) potassium chlorate ($KClO_3$) decomposing to form potassium chloride and oxygen gas.
(g) sodium sulphide reacting with iron (III) chloride to form sodium chloride and iron (III) sulphide.
(h) gasoline (C_8H_{18}) combining with oxygen to form carbon dioxide and water.
(i) the synthesis of tin (II) fluoride from the elements.
(j) the electrolysis of hydrochloric acid.

We have seen that the conductivity of solutions can be explained in terms of charged particles called ions. The question that sometimes arises at this point is, are ions present in all conducting solutions all the time or only when the electricity is turned on?

The following experiment might help to answer this question. There are two things to remember while you perform this experiment.

1. Regardless of which pair of reactants you choose to mix, *if* there is a chemical reaction it is always the same chemical reaction.
2. The product of the reaction you see is a binary compound.

3.11

Reactions Involving Solutions

test tubes
0.1 M potassium sulphide solution
0.1 M sodium sulphide solution
0.1 M potassium chloride solution
0.1 M sodium bromide solution
0.1 M copper (II) chloride solution
0.1 M copper (II) bromide solution

Warning

This experiment should be performed with the windows open.

1. Mix 5 ml of 0.1 M potassium sulphide solution with 5 ml of 0.1 M copper (II) bromide solution.
2. Mix 5 ml samples of any other pairs of reactants that you wish to try.

Questions

1. What substance do you think was formed in the first reaction?
2. What substances appear *not* to have reacted in the first reaction?
3. Write the equation for the first reaction showing *only* those things that reacted.
4. What ion(s) must be present in a
   (a) potassium sulphide solution;
   (b) copper (II) bromide solution;
   (c) sodium sulphide solution;
   (d) copper (II) chloride solution;
   (e) potassium chloride solution?

If we consider this experiment for a moment the following points may become obvious to you. Only those pair of solutions that contained *both* copper and sulphur reacted. Since the product must be a binary compound, it must be copper sulphide.

Using the valence of +2 for copper the formula becomes

CuS

If we go back to the first reaction

$K_2S + CuBr_2 \longrightarrow CuS + ?$

Something must be left over. The something must be potassium and bromine. We have no evidence to show that these two things react together so we leave them out of the equation. Consider potassium sulphide for a moment. The valence numbers tell us that in potassium sulphide potassium is +1 and sulphur is −2. In copper (II) bromide, copper is +2 and bromine is −1. In copper sulphide copper is +2 and sulphur is −2.

Since the copper (II) sulphide forms a precipitate, it must mean that both the copper (II) and the sulphur (−2) are able to separate from their original partners.

$$K_2S \longrightarrow 2K^{+1} + S^{-2}$$
$$CuBr_2 \longrightarrow Cu^{+2} + 2Br^{-1}$$

This splitting-up process is called *dissociation*. Some people refer to this as *ionization*. If we try to write the equation for the first reaction using the ionized solutions we get,

$$2K^{+1} + S^{-2} + Cu^{+2} + 2Br^{-1} \longrightarrow CuS + 2K^{+} + 2Br^{-1}$$

Remember, we have no evidence that potassium and bromine combine. If we now eliminate those ions that have not reacted in any way we get,

$$S^{-2} + Cu^{+2} \longrightarrow CuS$$

The K^{+1} and Br^{-1} ions merely stood around and watched while the Cu^{+2} and S^{-2} ions did all the reacting. For this reason, in this reaction, K^{+1} and Br^{-1} are called *spectator ions*. The equation we just wrote with the spectator ions removed is called the *net ionic equation*.

Apparently then, the ions must be present in conducting solutions all the time and not just when the current is turned on.

3.12

The Electrolysis of Aqueous Copper (II) Chloride

This lab exercise will require you to make some predictions based on what you have learned in this unit. You will be given the apparatus needed to electrolyze aqueous copper (II) chloride. Answer these questions *before* you begin the experiment.

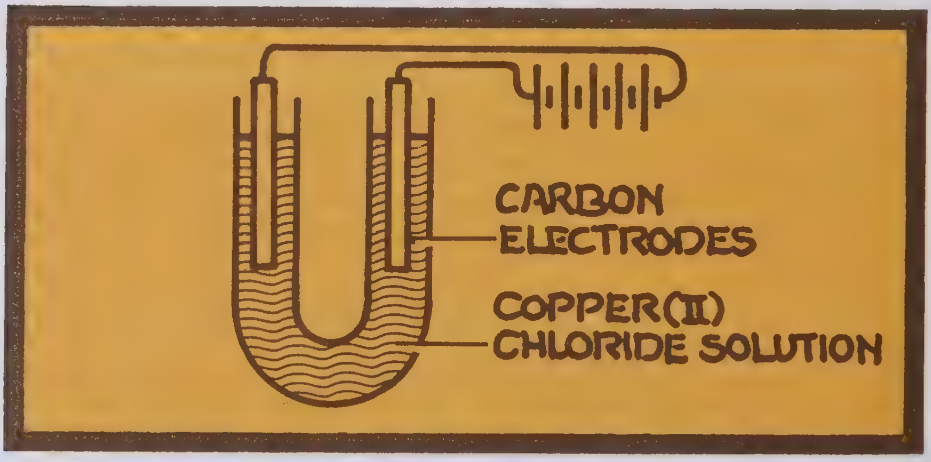
1. What ions will be present in the copper (II) chloride solution?
2. What migrations will take place when the current is turned on?
3. What observations will be expected at the cathode? at the anode?
4. Write equations for the reaction taking place (a) at the cathode, and (b) at the anode.

Now do the experiment.

0.5 M copper (II) chloride
2 carbon electrodes
U-tube
power source (6 V)
electrical leads with alligator clips

Electroplating of a car. Metal surfaces are often prepared to take paint and to resist corrosion. Corrosion-resistant coatings can be applied either by dipping or by electroplating. *Courtesy of Ford Motor Company of Canada, Ltd.*

1. Fill the U-tube to within 2 cm of the top with copper (II) chloride solution.
2. Place a carbon electrode in each arm of the U-tube. Connect the electrodes to the power source.
3. Turn on the power supply and allow the apparatus to run for several minutes. Record your observations carefully. Check the odour of the gas produced carefully (see Laboratory Skills, Unit 8).



Questions

1. How do your answers to the first set of questions look now that you have finished the exercise? Did your predicted observations come true?

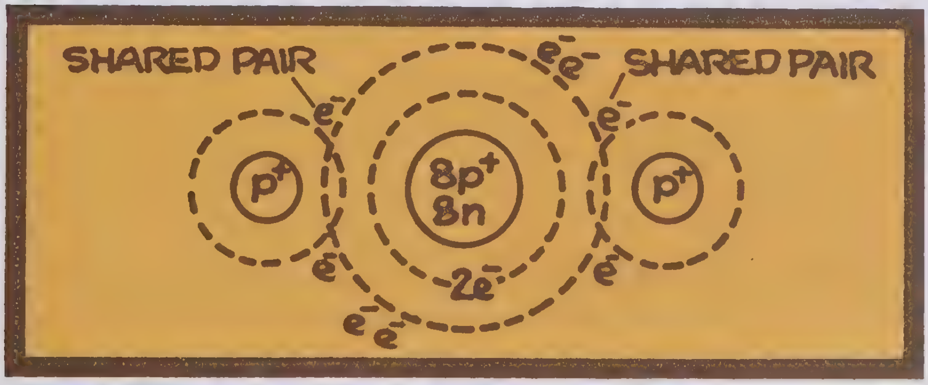
Types of Bonding

Hopefully you have spotted a tiny flaw in all of this. If you go back to Experiment 3.9 where we tested the conductivity of several substances we saw that some conducted well, some not so well, some poorly and some not at all. This idea of electrons transferring from one atom to another is fine for those substances that are electrolytes and conduct electricity, but what about those that do not? What about water itself?

When hydrogen reacts with oxygen we know we get water: $2H_2 + O_2 \longrightarrow 2H_2O$. But water doesn't conduct, at least not enough to make the bulb light up. How do we explain that?

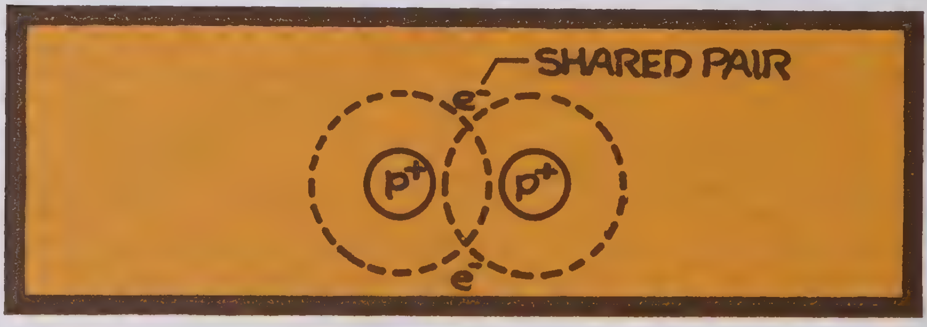
Well, in the first place, water does conduct a little bit. It is hard to detect but it is there. Actually, we think that something a bit different happens in water than we described earlier in the ionic compounds.

Oxygen wants to take an electron from the two hydrogens, but the hydrogen does not particularly want to give them up. Instead they do what any reasonable group would do when two members want the same thing. They share it. The two atoms move close enough together so their outer shells can overlap and the electrons in question move in *both* shells—an amicable solution of cooperation. In return for sharing its electron with oxygen, hydrogen gets the use of one of oxygen's electrons from time to time. Both atoms have full outer shells—part of the time. The molecule might look like this:



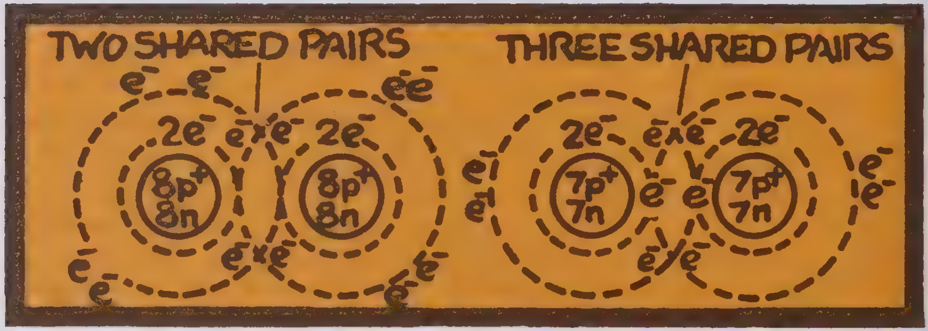
This molecule, held together by the sharing of a pair of electrons, is said to be a *covalent* molecule. Since it does not ionize, it does not readily conduct. Many non-conductors are covalent compounds. Actually, all molecules are a little bit covalent from time to time, even the ionic ones, because both members of a molecule attract all the electrons involved.

When one partner is much stronger than the other, the sharing is rather one-sided and may even look to be very ionic. When both partners are perfectly matched—as with two hydrogen atoms—the sharing is more even. The hydrogen molecule would be represented as:



With an oxygen molecule two pair must be shared:

And with a nitrogen molecule three pair must be shared:



Notice that in sharing, both atoms have the necessary number of electrons for stability *part* of the time.

Bonding then is caused by a competition for electrons. *All* nuclei attract electrons — some more than others. Some attract other atoms' electrons very strongly; some hold on to their own outer electrons weakly. Depending on the difference in strength involved, an outer electron may be transferred or shared with another atom. The resulting bond may be very ionic if transfer is complete or very covalent if sharing is perfect. It may also be somewhere in between.

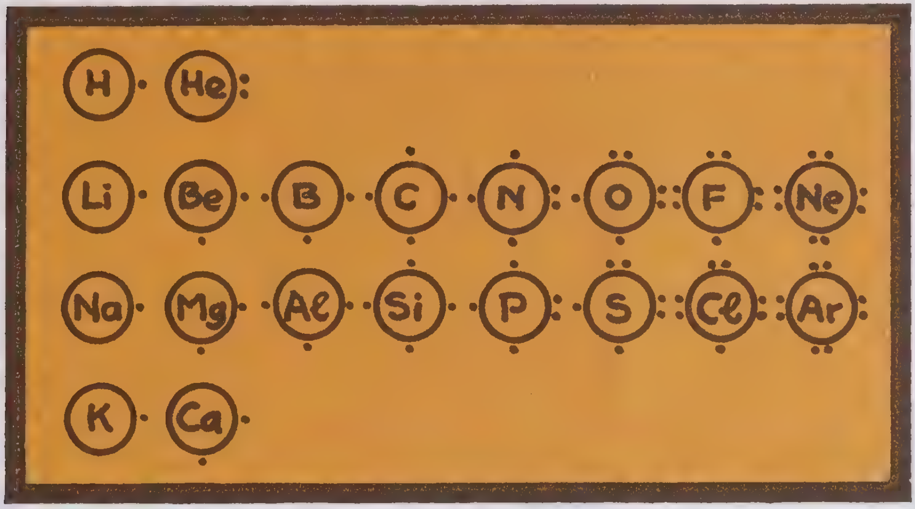
Table 3.10

All Ionic	*Mostly Ionic*	*In Between*	*Mostly Covalent*	*All Covalent*
CsF	NaCl	HCl	H_2O	H_2

Lewis Diagrams

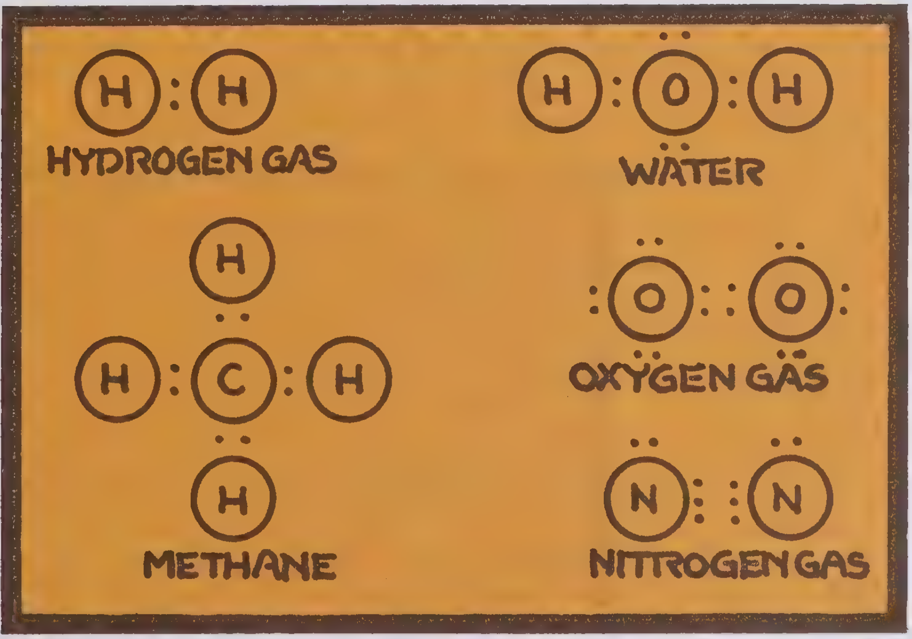
These diagrams are an attempt to show the electron arrangement in a molecule. They are rather cumbersome and hard to draw. We would like you to see two simpler versions of the same thing.

Since the only electrons involved in an exchange are found in the outermost occupied shell (the highest occupied energy level) some atomic diagrams look like this:


H
He
Li
Be
B
C
N
O
F
Ne
Na
Mg
Aℓ
Si
P
S
Cℓ
Ar
K
Ca

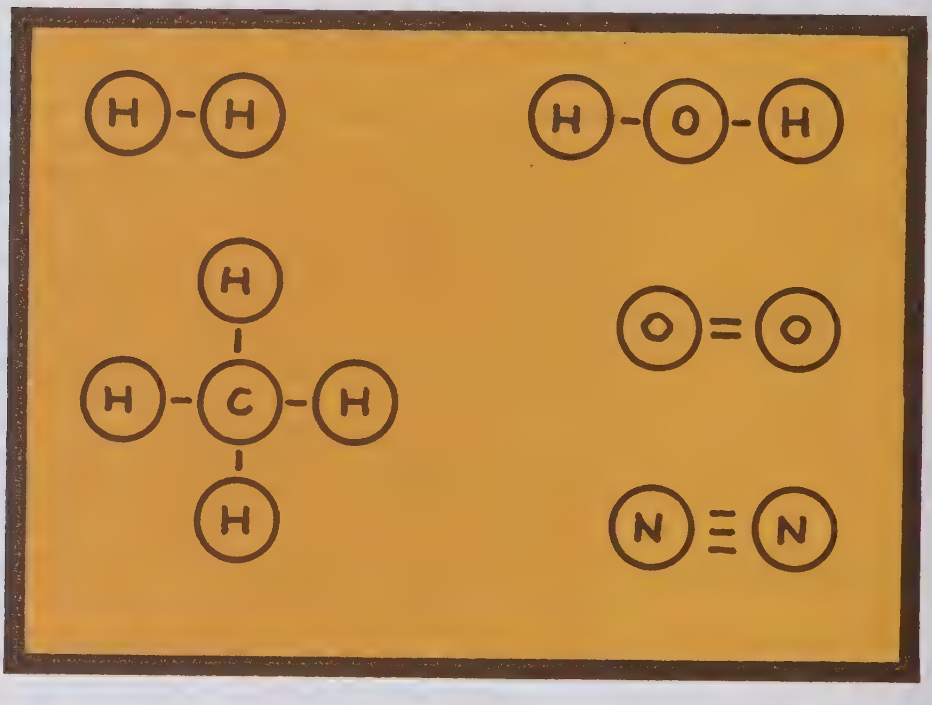
These diagrams are called *Lewis dot diagrams*. Each dot represents a valence shell electron. Notice that *only* the valence shell electrons are shown in the diagram.

Using Lewis dot diagrams we can represent molecules quite easily. Consider the following diagrams carefully:


H
H
HYDROGEN GAS
H
O
H
WATER
H
H
C
H
H
METHANE
O
O
OXYGEN GAS
N
N
NITROGEN GAS

The shared pairs of electrons are positioned between the nuclei of the atoms involved in the bond.

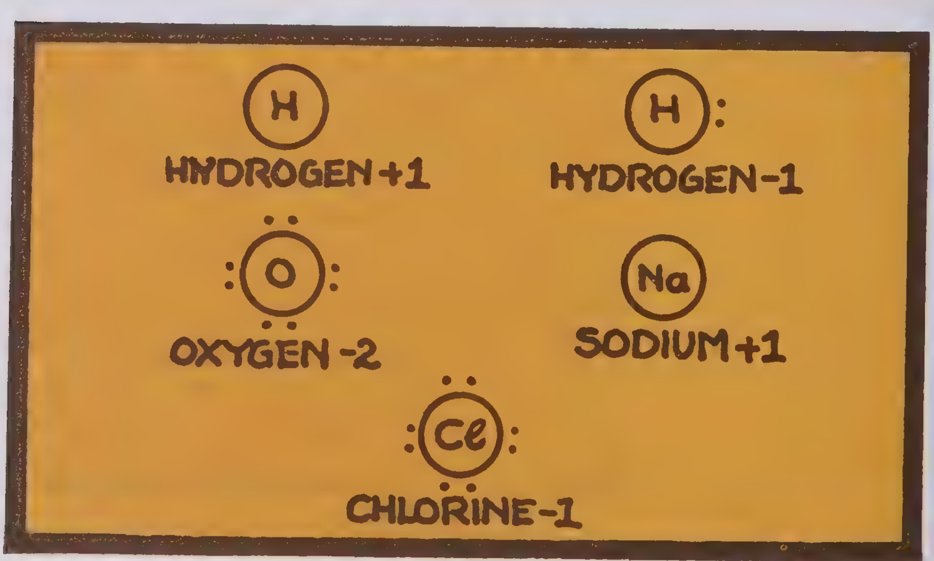
The next simplification simply substitutes a — for each shared pair of electrons. All other electrons are omitted from the diagram. Here are the same examples shown with this notation:


H-H
H-O-H
H
H-C-H
H
O=O
N≡N

As you can see, it would be quite simple to build models of these molecules using styrofoam spheres and toothpicks (or some other similar materials).

Generally this last notation is used only for substances that tend to be covalent in nature. The — is used to represent a shared pair of electrons and not one that is transferred as in ionic bonds.

Lewis dot diagrams may be used to show the electron configuration of ions. Here are some examples:


H
HYDROGEN +1
H:
HYDROGEN -1
O
OXYGEN -2
Na
SODIUM +1
Cl
CHLORINE -1

Problems

1. For each of the substances listed below carry out the following operations:
    1. Determine a possible formula.
    2. Draw the Lewis dot diagram of the molecule.
    3. Build a model of the molecule from the materials provided.
    4. Show each model to your teacher before you go on to the next.
        - (a) phosphorus (III) iodide
        - (b) carbon tetrachloride
        - (c) silicon carbide
2. Solid sulphur forms into an eight-membered ring with the formula S_8. Try to draw a Lewis dot diagram of this molecule.
3. Solid phosphorus forms into a molecule with the formula P_4. Try to build a molecular model of the phosphorus molecule.
4. The substance ethene has the formula C_2H_4. Suggest a possible structure for this substance.

Isotopes

There is one other thing that must be added before this unit is complete. As we have seen, the atomic number of an element is fixed by the number of protons in the nucleus. Change the number of protons and the element itself is changed.

There is no such fundamental change when the number of neutrons is changed. In fact there is no rule that enables us to determine the number of neutrons in an atom. As a result all atoms exist in several different forms called *isotopes* which have different numbers of neutrons. For example, hydrogen has one proton. It can exist in three different forms:

a single proton

a proton and a neutron

a proton and 2 neutrons

Since a proton has a mass of 1 u and a neutron has a mass of very nearly 1 u, the masses of these atoms would be:

1 u

2 u

3 u

These three isotopes of hydrogen all have the same atomic number, 1, but their atomic masses are different because they contain different numbers of neutrons.

Heavy water plant, Douglas Point, Ontario. Deuterium oxide (heavy water) is separated from ordinary water in plants such as this one. The heavy water is used in nuclear generating stations to moderate or control the reaction. *Courtesy of Ontario Hydro.*

If water is made from the first type of hydrogen the total mass of the water molecule will be:

H_2O
$1 + 1 + 16 = 18\ \text{u}$

If water is made from the second type of hydrogen:

H_2O
$2 + 2 + 16 = 20\ \text{u}$

This is the well-known heavy water used in nuclear reactors.

If water is made from the third type of hydrogen:

H_2O
$3 + 3 + 16 = 22\ \text{u}$

Chemists have given this the very technical name *heavy heavy water*.

Naturally, it is necessary to be able to identify the isotope we are dealing with. This is done by showing the atomic mass and the atomic number of the element along with its symbol in the following manner.

$${}^{1}_{1}H$$
$${}^{2}_{1}H$$
$${}^{3}_{1}H$$

These three isotopes of hydrogen are called hydrogen, ${}^{1}_{1}H$, deuterium, ${}^{2}_{1}H$ and tritium, ${}^{3}_{1}H$.

All that remains to be said is that deuterium and tritium are very rare. Only about one hydrogen atom in every 10 million is a deuterium atom. Even fewer are tritium. However, if we were to find the mass of 6.023×10^{23} hydrogen atoms we would find that they had a mass of just slightly more than the 1.00000 g we expected because of the presence of a few deuterium and tritium atoms. That is why the atomic mass of hydrogen is frequently listed as 1.008. This represents the average mass of 1 mol of hydrogen atoms as found in nature.

Chlorine exists in two principal forms,

$\text{Cl } 17^{35}$ and $\text{Cl } 17^{37}$

The published atomic mass of 35.45 indicates that only about one chlorine atom in four is $\text{Cl } 17^{37}$. The other three are $\text{Cl } 17^{35}$.

$$\begin{array}{rcl} 3 \times 35 & = & 105 \\ 1 \times 37 & = & \underline{\ 37} \\ \text{total} & = & 142 \\ \text{average} & & 142 \div 4 = 35.5 \end{array}$$

The most significant thing about isotopes is that some are more radioactive than others. As man learns to harness radioactivity and its associated atomic energy, isotopes will become more and more important.

Radioactivity

Merlin, the magician of King Arthur's mythical court, is thought to have been an alchemist. These old-time chemists spent their lives looking for either the Elixir of Life or the Philosophers' Stone.

The Elixir of Life was a liquid that was supposed to give eternal youth. You may recall Ponce de León and his quest for the Fountain of Youth in Florida. Mankind has long sought the life-prolonging Elixir.

The Philosophers' Stone has a more practical use. When mixed with another substance the Philosophers' Stone was supposed to turn the other substance into gold. The conversion of one element into another is called *transmutation*. From your knowledge of atomic structure you should recognize that transmutation would require an alteration of the positive charge in the nucleus.

As we have already mentioned, it was Henri Becquerel who discovered the phenomenon that we now call radioactivity. Becquerel was a physicist and so he asked a chemist, Marie Sklodawska Curie, to attempt to isolate the substance that produced the radioactivity. Marie Curie isolated two such substances; one called polonium and one radium.

Marie Curie

Wilhelm Konrad Röentgen

Otto Hahn and Lisa Meitner at the dedication of the new Hahn-Meitner Institute in Berlin, March, 1959

Ernest Rutherford was able to show that some radioactive substances produced alpha particles, which he correctly identified as He ions. Rutherford and his co-worker, Frederick Soddy, went on to show that atoms were not indivisible but that radioactive substances were constantly changing or transmutating into other elements. The alchemists had been partly correct. One element could change into another.

Uranium, ${}^{238}_{92}U$ has 92 protons.

Uranium atoms may eject an alpha particle, ${}^{4}_{2}He$, which contains two protons. This means that the original uranium nucleus has lost two protons and now contains only 90 protons. It has also had its mass reduced by 4, from 238 to 234. The element which is atomic number 90 is thorium. Uranium has become thorium.

$$ {}^{238}_{92}U \longrightarrow He^{++} + {}^{234}_{90}Th $$

We must not forget that ${}^{238}_{92}U$ has 92 electrons, whereas ${}^{234}_{90}Th$ has only 90 electrons. Two electrons are also emitted during the radioactive decay of uranium.

$$ {}^{238}_{92}U \longrightarrow H^{++} + {}^{234}_{90}Th + 2e^{-} $$

This type of radioactive decay is only one of several types. We could consider other types of decay, but it seems more important to look at other aspects of radioactivity.

Some substances are naturally radioactive. Radium is such a substance. As we observe a sample of ${}^{226}_{88}$ Ra we see it eject alpha particles to become radon, ${}^{222}_{86}Rn$.

$$ {}^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{4}_{2}He^{++} $$

This means that over an extended period some of the radium disappears as it changes to radon. If we were to start with 1000 g of

Pickering nuclear power station. At the nuclear power station at Pickering, Ontario, Ontario Hydro has harnessed the energy of nuclear transformations for peaceful purposes. *Courtesy of Ontario Hydro.*

radium, we would discover that 1620 years later we would have only 500 g of radium. 1620 years after that we would have 250 g of radium and 1620 years after that only 125 g of radium would remain. This means that regardless of how much radium we start with, in 1620 years only half will remain. This figure is called the *half-life* of radium. Every radioactive substance has a half-life which is different from all other radioactive substances. The half-life of $^{238}_{92}U$ is 4.5×10^9 years. The half-life of $^{222}_{86}Rn$ is only 3.8 days.

In a more extreme case, the half-life of $^{110}_{47}Ag$ is a mere 24 s.

All radioactive decay is an attempt on the part of the atom to become more stable. Any decay, therefore, results in the production of energy. This energy shows up as the kinetic energy of the ejected alpha particles and electrons or as gamma radiation. These are not the only forms of energy but they will do for our discussion now. The energy released by these transmutations is enormous.

So great was the interest shown in these reactions that scientists soon learned to create their own radioactive elements. It was Ernest Rutherford who first showed that if one element is bombarded with alpha particles the reverse of decay can take place. Rutherford showed that nitrogen could be turned to oxygen:

$$^{14}_{7}N + ^{4}_{2}He^{++} \longrightarrow ^{17}_{8}O + ^{1}_{1}H$$

Sometimes these bombardments produce a new isotope which is radioactive.

$$^{238}_{92}\mathrm{U} + {}^{4}_{2}\mathrm{He}^{++} \longrightarrow {}^{240}_{94}\mathrm{Pu} + 2\,^{1}_{0}\mathrm{n}$$

Plutonium is itself radioactive. This is called artificial radioactivity.

These new transmutations of small nuclei into larger ones have been carried to the ultimate with the production of elements that are never found naturally. Today we have many elements unknown in 1920. The periodic table has steadily grown from its old maximum of 92 elements until it now includes the trans-uranium elements, the elements larger than uranium; neptunium, plutonium, einsteinium and many more. We are confident that if we say there are 105 elements known today someone will manufacture the 106th tomorrow.

Radioactive elements, transmutation and artificial radioactivity will mean a great deal to the future of mankind. As we learn to control the processes of fission and fusion and harness the enormous energies involved, scientists hope to tap a new source of energy. Canadian scientists are world leaders in this area with the CANDU Reactor which has an excellent record of performance and safety.

The future of these radioactive materials is clear. As man learns more, the benefits will grow.

Table 3.11

Isotopes of Some Elements[a]

Carbon	$^{12}_{6}\mathrm{C}$	$^{14}_{6}\mathrm{C}$		
Barium	$^{131}_{56}\mathrm{Ba}$	$^{133}_{56}\mathrm{Ba}$	$^{137}_{56}\mathrm{Ba}$	
Silver	$^{107}_{47}\mathrm{Ag}$	$^{110}_{47}\mathrm{Ag}$	$^{111}_{47}\mathrm{Ag}$	
Calcium	$^{40}_{20}\mathrm{Ca}$	$^{41}_{20}\mathrm{Ca}$	$^{45}_{20}\mathrm{Ca}$	$^{47}_{20}\mathrm{Ca}$
Chlorine	$^{35}_{17}\mathrm{Cl}$	$^{36}_{17}\mathrm{Cl}$	$^{37}_{17}\mathrm{Cl}$	
Sulphur	$^{32}_{16}\mathrm{S}$	$^{35}_{16}\mathrm{S}$		
Uranium	$^{233}_{92}\mathrm{U}$	$^{234}_{92}\mathrm{U}$	$^{235}_{92}\mathrm{U}$	$^{238}_{92}\mathrm{U}$
Hydrogen	$^{1}_{1}\mathrm{H}$	$^{2}_{1}\mathrm{H}$	$^{3}_{1}\mathrm{H}$	

[a] Radioactive isotopes are in red.

Summary

No single list of discoveries could do justice to all the events involved in the growth of the atomic model for matter. Here is a chronological list of some of the important ideas and discoveries. Notice how the time scale is condensed as we near the present. There were more discoveries with less time between them. It seems that science is moving at an ever-increasing rate!

about 400 BC	Democritus and the Greeks
about 1645 AD	Gassendi revived the atomic theory and rekindled the spark left burning through the dark ages
1685	Newton supported atomism
1803	Dalton's experiments explained by atomic theory
1808	Gay-Lussac's Law of Combining Volumes
1811	Avogadro's Hypothesis
1858	Cannizzaro's experiments with atomic mass
1869	The Periodic Table by Mendeleev
1895	The first inert gas discovered by Ramsay and Rayleigh
1895	Röentgen discovered x-rays
1896	Becquerel discovered radioactivity
1897	Thomson's model of the atom based on electrons and protons
1898	The Curies discovered radium and polonium
1902	Ernest Rutherford proposed the theory of transmutation
1911	Ernest Rutherford proposed the nuclear atom
1913	Frederick Soddy discovered isotopes
1913	Niels Bohr proposed the shell model of the atom
1932	Chadwick discovered the neutron
1939	Otto Hahn discovered fission by neutron bombardment
1940	Enrico Fermi produced the first atomic pile

Somewhere around 1940 the brilliance of single personalities began to fade and research teams began to emerge. We find now that most major discoveries are the work of several scientists working together towards a common goal. It seems likely that this trend will continue. Don't be discouraged, however! There will always be a place for a clear thinking scientist who loves to solve problems. Mankind has always had more problems than solutions and that's a tendency not likely to change.

Questions and Problems

1. What evidence do we have for the existence of atoms?
2. What charge would be produced if:
   (a) a magnesium atom lost two electrons;
   (b) a chlorine −1 ion lost one electron;
   (c) a chlorine atom gained one electron?
3. Estimate the mass of a particle that contains 25 neutrons, 15 electrons and 13 protons.

4. Draw a diagram to represent Cl 17^{37}.
5. How many electrons could shell 7 hold?
6. What substances might be formed if an electric current were passed through a solution of copper (II) bromide?
7. Write a balanced formula equation to represent sodium combining with oxygen to form sodium oxide.
8. Draw a Lewis dot diagram to represent:
   (a) a chlorine ion;
   (b) a silicon tetrachloride molecule.

Some Words You Should Know

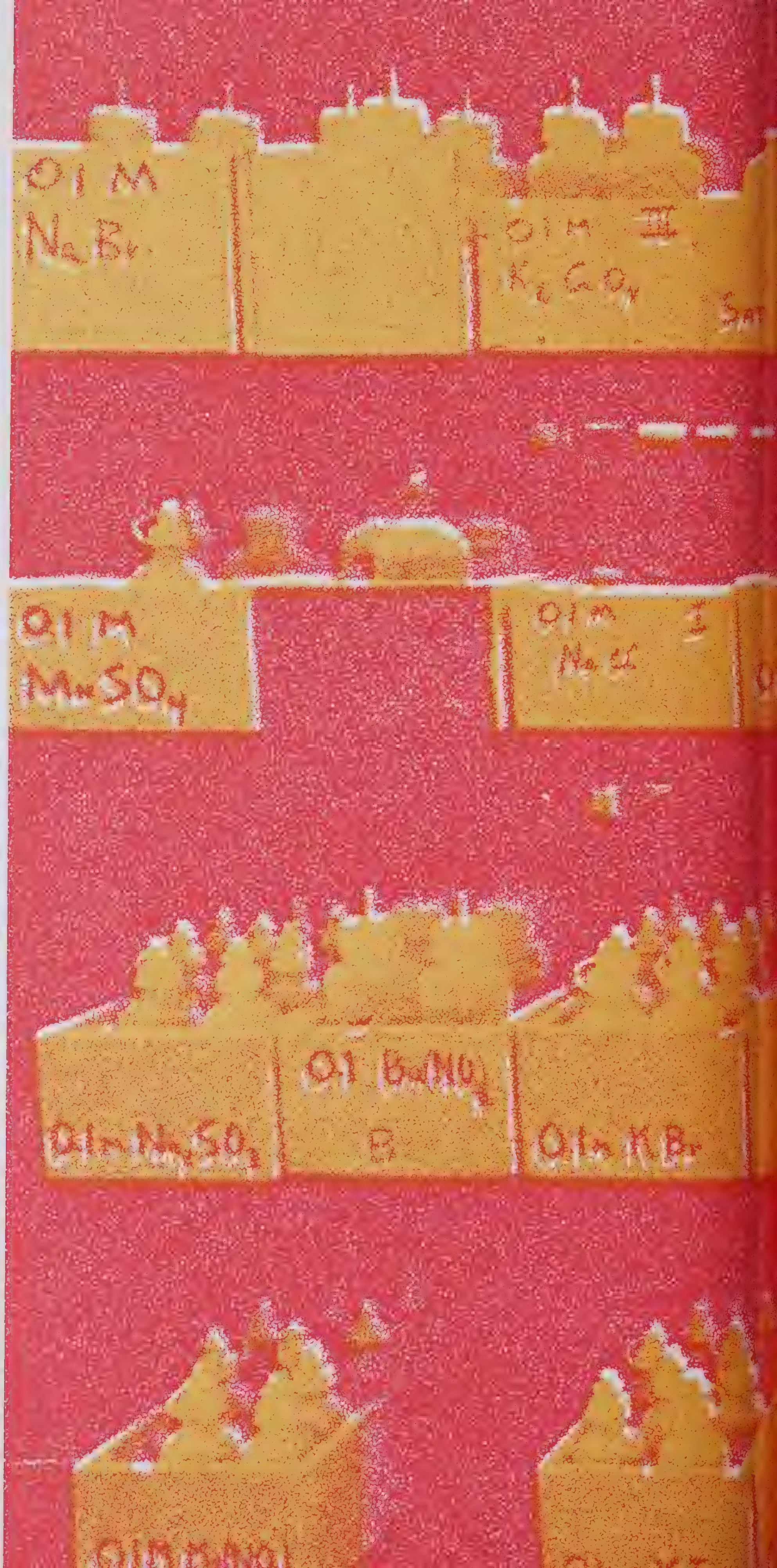
matter
molecule
atom
element
compound
energy
static electricity
charge
neutral
cathode ray
monatomic
diatomic
triatomic
conductivity
electrolysis
cathode
anode
electrode
word equation
formula equation
balanced equation
binary compound
planetary electron
electron
proton
neutron
sub-atomic particle
radioactive
alpha particle
beta particle
gamma ray
nucleus
orbit
energy level
principle quantum number
shared pair
covalent
bond
ionic bond
ion
dissociate
valence number
electrolyte
shell
spectrum
atomic number
atomic mass
Lewis dot diagram

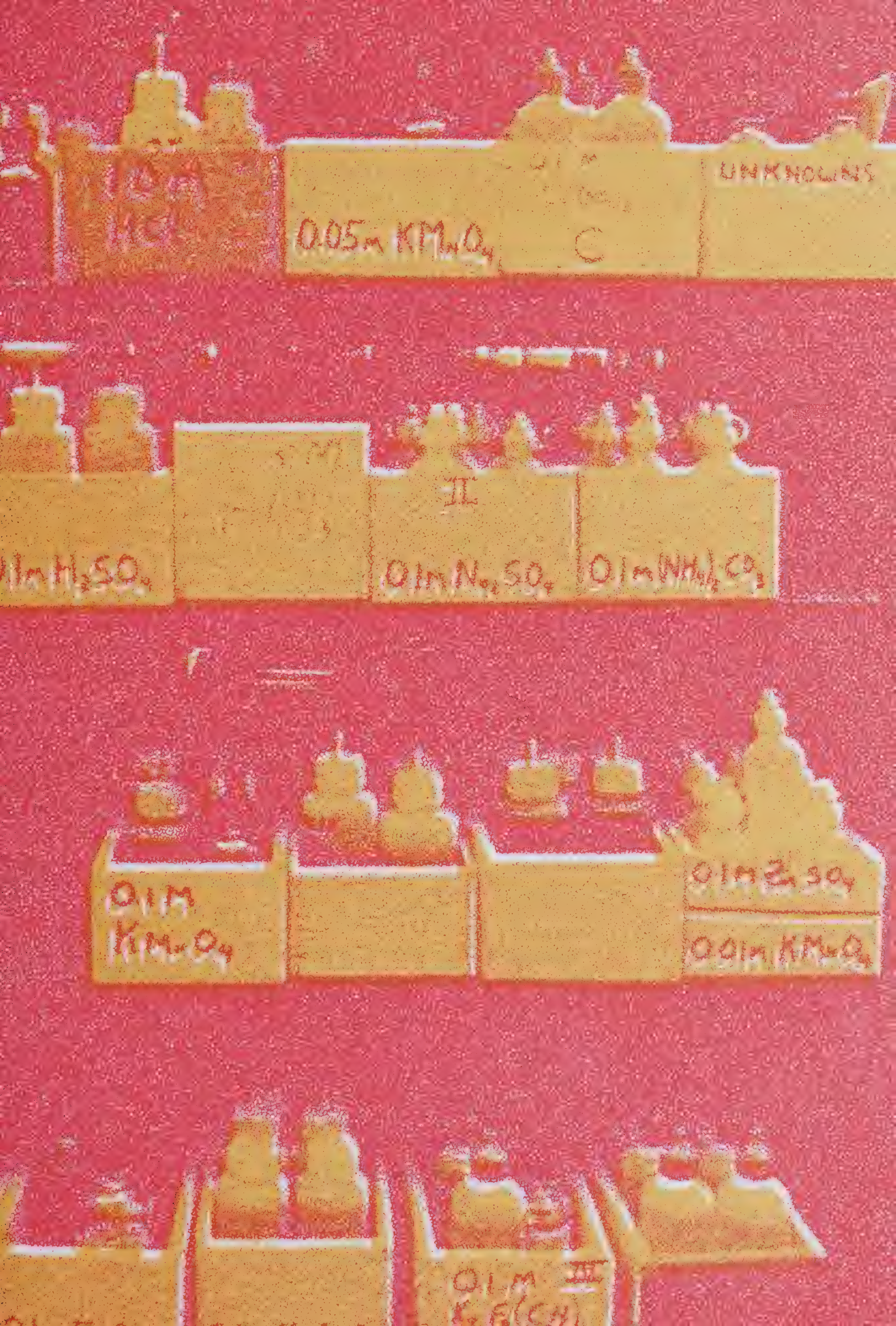
Some Names You Should Recognize

J.J. Thomson
Ernest Rutherford
James Chadwick
Hans Geiger
Ernest Marsden
Wolfgang Pauli
Neils Bohr
Henri Becquerel
Marie Curie
Pierre Curie
Robert Millikan
Max Planck
Svante Arrhenius
John Grey
Eugene Goldstein
William Crookes
Wilhelm Röentgen
Otto Hahn
Lise Meitner

4

Oxygen/Catalysis/
Acids and Bases/
Neutralization/
Molarity/Titrations/
Oxidation/
Reduction/
Stoichiometry/
Formula


0.1 M NaBr
0.1 M II K2CrO4
0.1 M Na2SO4
0.1 M I NaCl
0.1 m Na2SO3
0.1 Ba(NO3)2 B
0.1 m KBr


1.0 m HCl
0.05 m KMnO4
C
UNKNOWNS
0.1 m H2SO4
II
0.1 m Na2SO4
0.1 m (NH4)2CO3
0.1 m KMnO4
0.1 m ZnSO4
0.01 m KMnO4
0.1 m K3Fe(CN)6
III

Joseph Priestley

Joseph Priestley was born in England in 1733 and died at the age of 71 in 1804. As a boy he was extremely interested in science and continued his studies as a hobby as he matured. When he grew up he became a Unitarian minister and at the age of 27 he took over a parish in Leeds, England.

Strangely enough, a brewery happened to be next door to his church. With his interest in science the brewery provided him with an unlimited supply of carbon dioxide to carry out many experiments with the gas (Carbon dioxide, hydrogen and air were the only gases known at the time.)

Priestley collected the gas for his work by displacing water from a container. Accidently he happened to taste the water that was displaced and he found that carbon dioxide gas had given the water a rather tart taste. This water he called "seltzer." Today, we call it soda water. All he needed to do was add flavouring and sugar and he would have developed the soft drink.

Nevertheless, he concluded that the tart taste was caused by some of the gas dissolving in the water. As a result he decided to collect the gas over mercury rather than water. Because he turned to mercury he was able to isolate and study other gases as well. Although he did not use present-day names for these gases his writings lead us to believe that he isolated ammonia, hydrogen chloride and sulphur dioxide. All these gases are quite soluble in water.

Priestley's most important work came in 1774. He discovered that when mercury was heated in air it formed a brick red coating or "calx" as he called it. (Today we call this red compound mercuric oxide.)He managed to separate some of this calx from the mercury. A small portion was placed in a test tube which he heated with heat from the sun. He used a lens to concentrate the light rays onto the material in the test tube. The red calx disappeared and mercury droplets formed on the sides of the tube. A glowing splint lowered into the test tube burst into flame, something that just did not happen in air.

Burning at this time was misunderstood. It was believed that combustible materials were rich in a substance called phlogiston. When a substance burned, it released its phlogiston. Priestley was a firm believer in this theory even to the time of his death.

Priestley concluded that when the red mercury calx was heated it released air that was greatly lacking in phlogiston. When the splint was lowered into this air it accepted the phlogiston with unusual ease causing the splint to burst into flame. He called this gas dephlogistonated air.

Priestley collected large quantities of dephlogistonated air. He found that combustible materials burned more brilliantly in the gas, mice thrived and when he inhaled it himself he felt unusually "light" or "high."

Because Priestley was so thoroughly convinced of the phlogiston theory, he missed the real significance of his discovery. The gas that he had isolated was of course oxygen. Even though he did not name the gas as such he is credited with its discovery.

Gas Production

In our study of chemistry up until now we have gathered sufficient evidence to indicate that matter is made up of particles that we call molecules. The molecules in many cases may be broken down into smaller particles that we call atoms. This means of course that atoms can join with each other to form more complex particles. Further, we discovered that when particles of matter combine to form new particles certain laws govern their joining. Chemical bonds hold atoms together into molecules and the joining of particles is a chemical reaction. We are now going to study certain chemical reactions and at the same time determine some properties of a few of the more common elements and compounds.

The following experiments will result in the production of one of the most useful elements to mankind. Identify the element and determine some of its properties.

4.1

The Effect of Manganese Dioxide on Hydrogen Peroxide

test tube
splint
filter paper
balance
watch glass
funnel
3% hydrogen peroxide
manganese dioxide

Part A

1. Place about 3 ml of hydrogen perioxide into an 18 × 150 mm test tube.
2. Place a glowing splint into the mouth of the test tube. Observe.

3. Remove the splint and add a small amount of manganese dioxide to the test tube (just enough to cover the tip of your spatula). Immediately lower a glowing splint into the mouth of the test tube. Record all observations.

Questions

1. What happened to the glowing splint when it was lowered into the test tube above the hydrogen peroxide?
2. What happened in the test tube when the manganese dioxide was added?
3. What happened to the glowing splint? What does the glowing splint test indicate to us?

Masses must be very carefully determined in this experiment.

Part B

1. Carefully find the mass of a piece of filter paper. Record the mass of the paper.
2. Fold the paper into a filter cone. (Fold the paper first in half and then into quarters. Open the cone so that one side is three layers thick and the other is one layer thick.) Add about 1 g of manganese dioxide to the filter cone. It is not necessary to have exactly 1 g but the mass must be taken exactly. Record the mass of the manganese dioxide and the filter cone.
3. Place the filter cone with the manganese dioxide in a funnel. Pour 5 ml of hydrogen peroxide into the manganese dioxide and through the filter (add the peroxide carefully).
4. When most of the peroxide has passed through the filter remove it from the funnel, place it on a watch glass which is labelled with your name and set it aside until next day. It will dry out in the meantime.

The remaining procedure must be done on the next day.

5. Get your watch glass and find the mass of the dry filter cone and residue. Record the mass of filter cone and residue.
6. Scrape some of the residue from the filter paper. Add it to about 2 ml of hydrogen peroxide in a test tube. Immediately lower a glowing splint into the test tube. Record all observations.
7. Calculate the mass of manganese dioxide used (mass of manganese dioxide + filter cone − mass of filter cone).
8. Calculate the mass of the residue (mass of residue + filter cone − mass of filter cone).

Questions

1. How does the mass of the manganese dioxide used in Part B compare to the mass of the residue after the reaction? (Class results should be compared.)
2. What effect did the residue have on the peroxide in Step 6? What happened to the splint?
3. The manganese dioxide in this chemical reaction is a catalyst. The action of a catalyst is called *catalysis*. Define "catalyst" basing your definition on your observations.
4. What one word would best describe what has happened to the hydrogen peroxide?
5. Write a balanced chemical equation for the reaction of the hydrogen peroxide.

Peroxides

In the last experiment you were using two substances with slightly different names, hydrogen peroxide and manganese dioxide. *Peroxides* are a very select and very small class of chemical compounds. They have been defined in many ways but probably the simplest way to identify a peroxide is to look at the valence of the oxygen.

hydrogen peroxide — H_2O_2

Hydrogen can only be +1, so oxygen must be −1. Peroxides are substances that contain oxygen in its −1 valence state.

Other members of this group are:

potassium peroxide	K_2O_2
sodium peroxide	Na_2O_2
barium peroxide	BaO_2

That's it. There are only four that you need worry about!

Manganese *dioxide* is the old-fashioned name for manganese (IV) oxide.

There are lots of examples of the old system of naming that are still used.

carbon dioxide ———— carbon (IV) oxide
lead dioxide ———— lead (IV) oxide
water ———— hydrogen oxide
carbon tetrachloride ———— carbon (IV) chloride

About the only thing we can say is that old habits die hard and these names will probably be around for a long time. The prefixes used help to identify the substance:

mono — one
di — two
tri — three

tetra	four
penta	five
hexa	six
hepta	seven
octa	eight
nona	nine
deca	ten

These prefixes were inserted in the name. For example, copper sulphate pentahydrate was the name given to $CuSO_4\ 5H_2O$.

Here are some others:

carbon monoxide	CO
carbon dioxide	CO_2
phosphorus trioxide	P_2O_3
phosphorus pentoxide	P_2O_5
carbon tetrachloride	CCl_4
uranium hexafluoride	UF_6

Catalysis

Catalysts are substances that can alter the *rate* or speed of a chemical reaction. There are many catalysts in the world. One very common one is the substance added to liquid plastic to make it solidify. The liquid plastic is always in the process of solidifying, but without the catalyst it would take 12 to 16 months. With the catalyst it requires only about one hour.

A catalyst cannot change the outcome of a reaction. It can only change the rate of the reaction. We can think of many catalysts that speed up reactions but there are also catalysts that slow down reactions. They are frequently called *inhibitors*.

Perhaps the greatest concentration of catalysts is found inside your body. Biological catalysts are called *enzymes*. There are many enzymes present in a living organism. Try chewing a cracker without swallowing any of it for a minute or more. It will soon begin to taste quite sweet. This is the result of the catalyst present in saliva speeding up the breakdown of starch into sugar.

In the next experiment you will again see a catalyst at work.

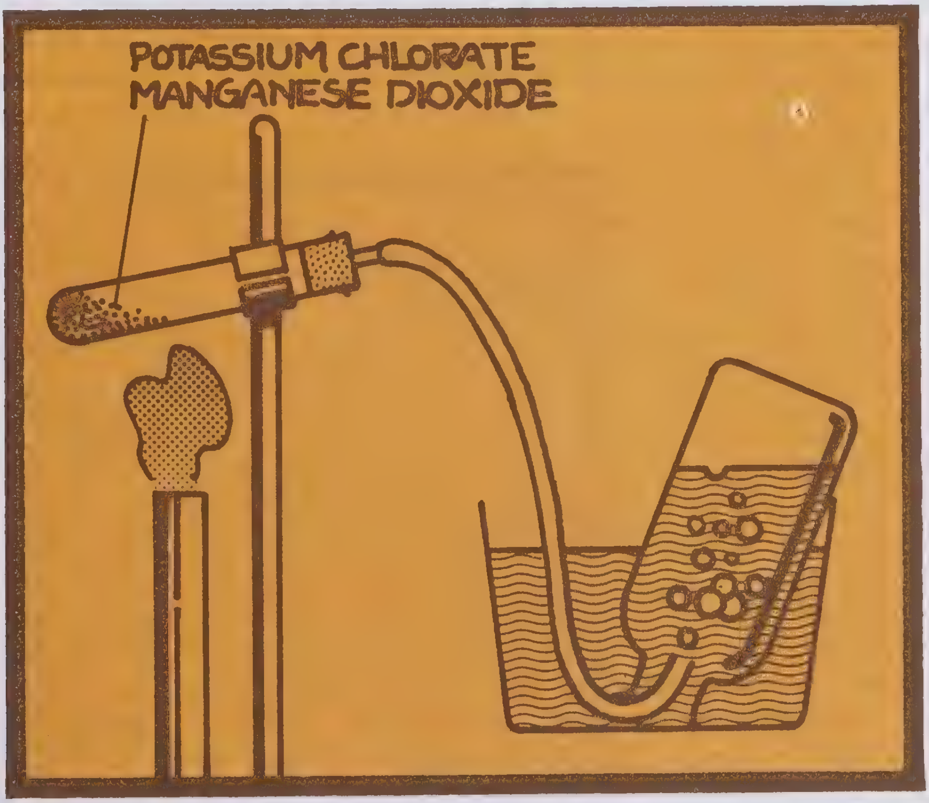
4.2

Properties of Oxygen

Warning

Extreme caution must be exercised while doing this experiment.

test tube
retort stand and clamp
stopper
rubber and glass tubing
pneumatic trough
4 gas bottles
bunsen burner
4 glass plates
splint
deflagrating spoon
asbestos liner
crucible tongs
potassium chlorate
manganese dioxide
charcoal
sulphur
magnesium ribbon
sodium
red and blue litmus
bromthymol blue solution



1. Prepare a mixture of five parts potassium chlorate and one part manganese dioxide. Place about 3 cm of the mixture along the side of an 18 × 150 mm test tube (see diagram). Assemble the remaining apparatus as shown in the diagram. The gas bottles (four in all) are first filled with water, then inverted into the tray of water.

Warning

Do not place the delivery tube in the water until the mixture in the test tube is bubbling.

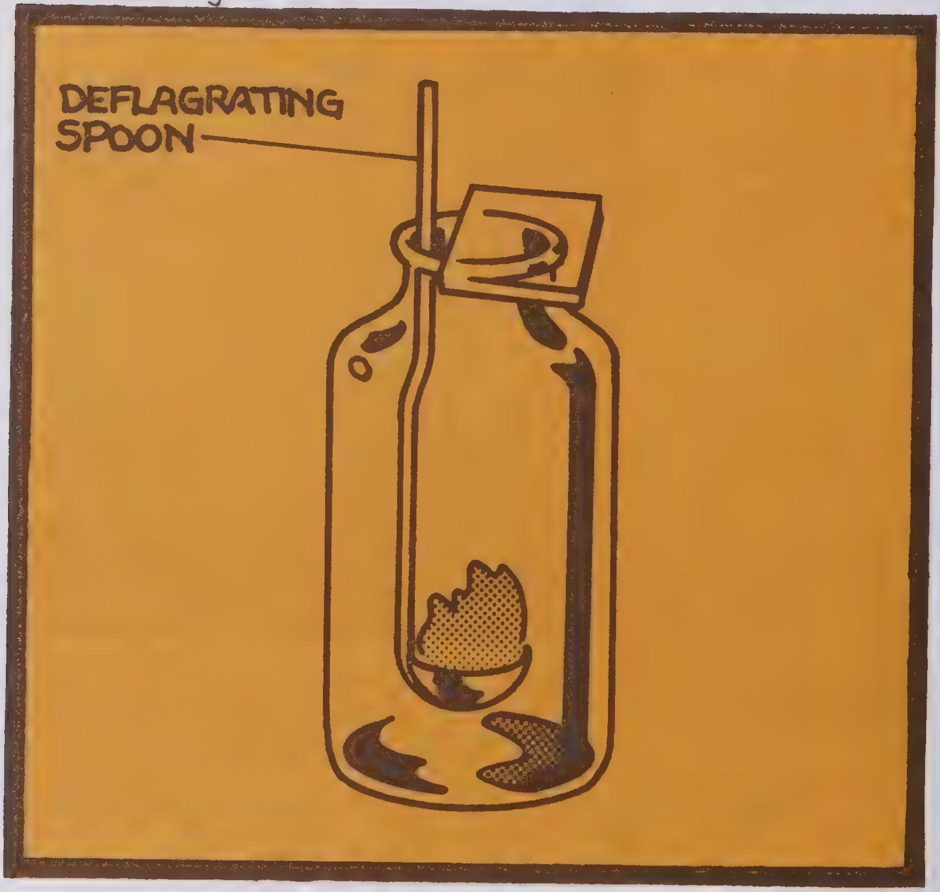
2. Begin heating the mixture in the test tube. When the mixture is bubbling place the open end of the delivery tube into the tray of water and under the mouth of one of the gas jars.

Fill each bottle in turn with the gas. When the last bottle is filled remove the delivery tube from the water then remove the flame from the test tube.

Warning

Remove the heat from the test tube *after* removing the delivery tube from the water.

3. Place a glass plate over the mouth of each jar of gas then turn each upright on the desk.
4. Remove the lid from the first of the four jars and insert a glowing splint into the jar. What gas have you produced?
5. Put about 10 ml of water into the bottom of the three remaining bottles of gas.



6. Line a deflagrating spoon with asbestos then put some small lumps of charcoal in it. Ignite the charcoal in your bunsen flame then lower the charcoal into a bottle of gas as shown in the diagram. When burning ceases remove the spoon, cover the bottle and shake well.
7. Add small pieces of red and blue litmus paper to the liquid in the jar. Observe. Add a few drops of neutral (green) bromthymol blue solution. Record all observations.
8. Repeat Steps 6 and 7, only use sulphur instead of charcoal.

Warning

Do not look directly at the burning magnesium.

9. Hold a 5 cm strip of magnesium ribbon with your crucible tongs. Ignite the magnesium ribbon then lower it into the remaining bottle of oxygen. When burning ceases allow the white ash to drop into the water. Shake the bottle then add the litmus paper and bromthymol as in Step 7.
10. Your teacher will repeat Steps 6 and 7 using a lump of sodium metal as a demonstration.

Questions

1. What gas is produced when potassium chlorate is heated? How do you know?
2. The manganese dioxide is once again a catalyst. Write a balanced chemical equation for the decomposition of potassium chlorate.
3. Elements burn in oxygen to produce their oxides. Write balanced chemical equations for the burning of magnesium, sodium, carbon and sulphur in oxygen.
4. Litmus paper and bromthymol blue are acid-base indicators. Litmus is red in acid, blue in base. Bromthymol is yellow in acid, blue in base. What type of solution was produced when the product of the combustion of carbon was dissolved in water? of sulphur? State the observations upon which you base your conclusions.
5. Both carbon and sulphur are non-metals. It appears that the oxides of non-metals dissolve in water to form what type of solution (acidic or basic)?
6. What type of solution was produced when the product of combustion of magnesium was dissolved in water? of sodium? State the observations upon which you based your conclusions.
7. Both magnesium and sodium are metals. It appears that the oxides of metals dissolve in water to form what type of solution (acidic or basic)?

8. Draw up a table similar to Table 4.1 and summarize the properties of the products of the combustion of the four elements and the solutions of the combustion products.

Table 4.1

Product of Combustion	*Carbon*	*Sulphur*	*Magnesium*	*Sodium*
Physical state				
Distinctive physical properties				
Name of oxide				
Solution of Oxide				
Effect on litmus				
Effect on bromthymol blue				
Type of solution				
Formula of acid or base produced				

Percentage Composition of Oxygen in Air

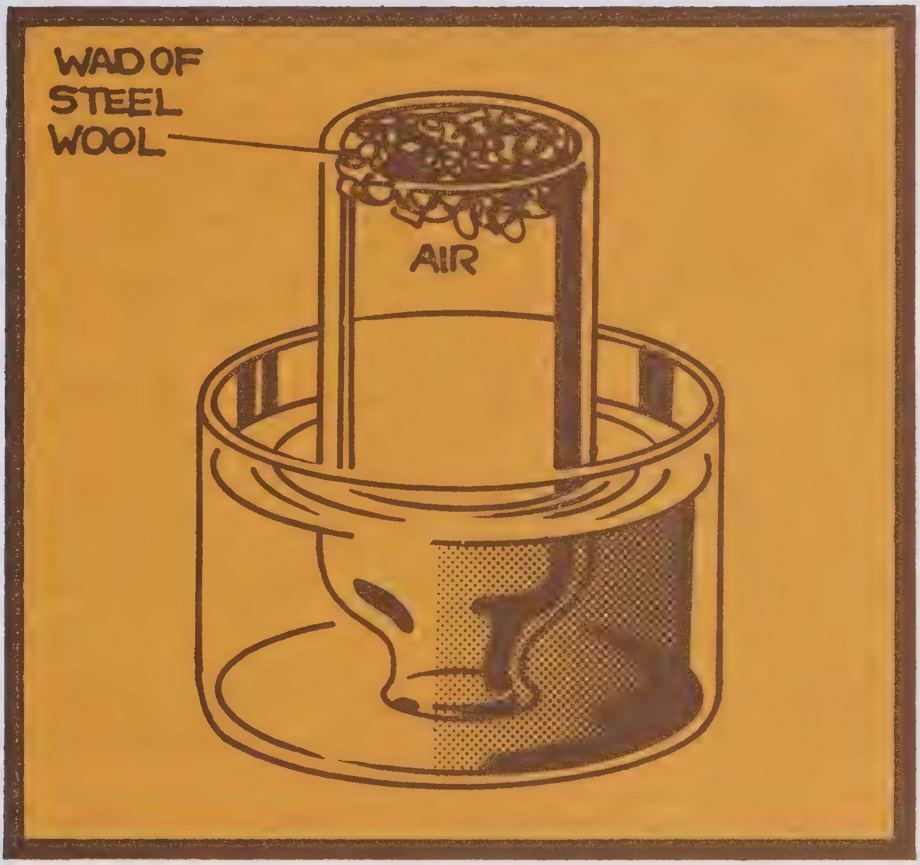
The next experiment is divided into two parts. In Part A you will be trying to determine the proportion of oxygen in air. In Part B you will be trying to determine how much oxygen can be produced by a given amount of potassium chlorate. You should do Part A first until you get to the end of Step 3, at which time you can go ahead and do Part B.

4.3

The Proportion by Volume of Oxygen in the Air

Part A

beaker
gas jar
large graduated cylinder
tray
glass plate
splint
steel wool
carbon tetrachloride



1. Remove the oil film from a wad of steel wool by rinsing it several times in some carbon tetrachloride in a beaker. (This procedure should be done in a fume cabinet.)
2. When the wool has dried push it to the bottom of a gas jar. Determine the volume of the remaining space in the jar by first filling the jar with water then pouring it out into a graduated cylinder. Record the volume of the gas jar.
3. Invert the jar into a tray that is half filled with water. Set it aside and leave it for a two-day period.
4. Cover the mouth of the jar with a glass plate then turn the jar upright. Lower a burning splint into the jar. Pour the water that is in the jar into a graduated cylinder. Record the volume of water that entered the jar.
5. Remove the steel wool and examine it carefully.

Questions

1. Describe and account for the appearance of the steel wool after the two-day period.
2. Has *all* of the steel wool changed?
3. Explain why the burning splint behaved as it did.
4. Why did water enter the jar as the reaction proceeded?
5. What volume of air was in the jar to start?

6. What does the volume of water that entered the jar represent?
7. From your data calculate the percentage of oxygen in the air.

Volume of oxygen/volume of air × 100 = percentage oxygen in the air

Part B

test tube
retort stand and clamp
stopper
rubber and glass tubing
pneumatic trough
gas bottles
burner
glass plates

1. Find the mass of a test tube and put about 5 g of potassium chlorate into it. Record accurately the mass of the potassium chlorate used.
2. Assemble the apparatus as in Experiment 4.2.
3. This time place the delivery tube inside the collecting bottle as soon as the *first signs* of bubbling appear.
4. Continue heating the potassium chlorate until it has completely solidified. If you find it necessary to change collection bottles try to do it as quickly as possible.
5. Remove the delivery tube and the burner. Allow the test tube to cool.
6. Determine the volume of gas collected.
7. Find the mass of the test tube and contents and determine the mass lost by the potassium chlorate.

Questions

1. In order to compare class results we will base our calculations on 1 mol of potassium chlorate. What is the mass of 1 mol of potassium chlorate, $KClO_3$?
2. If you had taken exactly 1 mol of potassium chlorate instead of approximately 5 g, what volume of oxygen would you have collected?
3. What was the mass of the oxygen you collected?
4. How many moles of oxygen did you collect?
5. From your experiment, what is the mass of 1 mol of oxygen?
6. How does this figure compare with the figure you obtained in the gas bag experiment?

7. How does this figure compare with the known gram molecular mass of oxygen?
8. How many moles of oxygen were produced by each mole of potassium chlorate in this experiment?
9. How many moles of oxygen would you expect to be produced by 1 mol of potassium chlorate?
10. Give some possible sources of error in this experiment.
11. Give some conditions that matter in this experiment.

Acids and Bases

In our discussion of oxygen and its properties you have seen the words acid or acidic and base or basic. You have surely been asking yourself what we mean by an acid or a base. The following experiments will show us what acids and bases are and how they behave.

4.4 Properties of Acids

4 test tubes
glass rod
splint

0.1 M hydrochloric acid
0.1 M nitric acid
0.1 M acetic acid
0.1 M sulphuric acid
red and blue litmus
phenolphthalein solution
bromthymol blue solution
magnesium ribbon
sodium carbonate

1. Set up four test tubes each containing 3 ml of hydrochloric acid.
2. Test the solution in the tube with red and blue litmus paper.
3. Draw a few drops from the test tube with a glass rod and place them between your thumb and index finger. Rub your thumb and finger together to note the feel. Taste a small drop of the acid. Add a few drops of pink penolphthalein solution to the remainder of the acid in Tube 1.
4. Add a few drops of neutral bromthymol blue to Tube 2.
5. Add about 3 cm of clean magnesium ribbon to Tube 3.
6. Add a small amount (enough to cover the tip of your spatula) of sodium carbonate to Tube 4. Immediately insert a burning splint into the tube.
7. Repeat Steps 1 to 6 using each of the other three acids listed.
8. Summarize your results in tabular form as illustrated.

Table 4.2

Acid:	*Acetic*	*Nitric*	*Sulphuric*	*Hydrochloric*
Taste				
Feel				
Litmus				
Phenolphthalein				
Bromthymol				
Carbonate				
Magnesium				
Formula				

Questions

1. List all the properties that are common to these four acids.
2. How would you identify an acid?
3. Determine the formula for each acid used in this experiment.
4. What do you think it is that gives an acid its properties?

You may have some trouble with the last line of Table 4.2. It appears that for some of these substances more than one formula is possible. This seems to apply especially to acids.

Sulphur, for example, can form three different oxides depending upon the conditions present. These are:

$2S + O_2 \longrightarrow 2SO$ sulphur (II) oxide

$S + O_2 \longrightarrow SO_2$ sulphur (IV) oxide

$2S + 3O_2 \longrightarrow 2SO_3$ sulphur (VI) oxide

Under normal conditions sulphur (IV) oxide, SO_2, is the most abundant product. When this oxide dissolves in water H_2SO_3, sulphurous acid is produced. $SO_2 + H_2O \longrightarrow H_2SO_3$

We will expand somewhat on this idea a little later.

When you defined an acid in terms of its properties you were, in effect, describing how to identify an acid. You were describing or defining an acid according to its taste, feel, effect on various indicators and so on. Such a definition is called an *operational* definition.

The operations you must perform to identify an acid are taste, smell, touch and effect on indicators.

On the other hand, if you define an acid as a substance that contains hydrogen ions, then you are not describing *how* to identify an acid, but *why* it is an acid. Such a definition is called a *conceptual* definition.

For example, suppose you were asked to define a tree.

A tree is a plant with a woody stem. It has bark on the outside of the stem. It has well-developed roots. It has either leaves or needles. This is an *operational* definition. It tells me what to look for when I try to identify a tree.

On the other hand we could define a tree as a plant organism that derives its energy from the reaction between carbon dioxide and water in the green tissues of its cells. This is a *conceptual* definition. (Which any good biologist would cut to pieces.)

Give an operational and conceptual definition for each of the following:

1. a circle;
2. a ghost;
3. a pocket calculator.

4.5 Properties of Bases

sodium hydroxide
potassium hydroxide
ammonium hydroxide
saturated barium hydroxide

1. Repeat all of the procedures from Experiment 4.4 with each of the four bases listed above. Once again summarize your results in tabular form just as you did for the acids.
2. List *all* of the properties that were common to each of the bases used.

Questions

1. Give an operational definition of a base.
2. Give a conceptual definition of a base.

The properties of acids and bases should not be entirely new to you. You have encountered examples of these substances most of your life. To illustrate this point do the following experiment.

4.6 Common Acids and Bases

indicators
liquids provided by the student

1. Get from your home samples of as many liquids as possible. Such things as milk, vinegar, fruit juices, cleaning fluids or soft drinks may be used. The samples need only be small.

Warning

Some of these liquids may be poisonous.

2. Carry out some of the tests from Experiment 4.4 on each of the solutions that you have provided. Remember—no tasting! Classify each as acid, base or neither acid nor base.
3. List on the blackboard the liquid and whether it was acid or base. Add to the list only those liquids that someone else has not yet provided.
4. Add the list to your notebook. Put acids in one column and bases in a second column.

Naming Acids and Bases

As in any area of knowledge, being able to speak the language is over half the battle. All vocations, all subject areas, even the arts and crafts have specialized vocabularies that only an insider knows. A large part of this course is vocabulary. If you can't speak the language, you won't understand what's being said.

Bases

Bases are usually fairly easy to name. As you have already seen, bases are formed when metallic oxides are dissolved in water. Bases are always called hydroxides. (That comes from hydro—water—and oxide—oxide.) To name a base, give the name of the metal it contains and follow it with hydroxide.

$NaOH$	sodium hydroxide
KOH	potassium hydroxide
$Ca(OH)_2$	calcium hydroxide

Since the OH always appears in a base it has been given the name *hydroxyl group*. It is a fairly good sign of a base if it appears in conjunction with a metal or some other positive ion.

Name these bases:

(a) $LiOH$
(b) $Al(OH)_3$
(c) $Zn(OH)_2$
(d) $Sr(OH)_2$
(e) $Mg(OH)_2$

Acids

Acids are a bit more complicated than bases because there seems to be more of them. Acids tend to come in two kinds.

Organic Acids

Many acids are found in foods, plants and animals. These are acids produced by living organisms and are called *organic acids*. They usually contain carbon, hydrogen and oxygen. Two examples are:

CH_3COOH —acetic acid or vinegar. This is produced by fermenting sugar.

HCOOH —formic acid. This is the substance injected into a bite by stinging ants. Formic acids used to be produced by distilling the bodies of stinging or biting ants and collecting the distillate.

Other examples of organic acids are:

(a) citric acid in citrus fruits;
(b) malaic acid in apples;
(c) lactic acid in sour milk.

There are literally thousands of organic acids and millions of organic compounds. In fact, there are so many organic compounds that there is a whole branch of chemistry called *organic chemistry* that deals only with these substances. We will be looking at some organic chemistry in Unit 6.

Mineral Acids

Acids which tend not to be found in living organisms are called *mineral acids*. They are usually formed when non-metallic oxides or binary compounds containing hydrogen are dissolved in water. Let's discuss the latter first.

Binary Acids

These acids contain only two elements, one of which is hydrogen. All these acids are formed when a binary gas dissolves in water. They are all identified by the prefix *hydro*. All binary acids have names ending in *ic*.

Gas	*Gas Formula*
hydrogen fluoride	HF
hydrogen chloride	HCl
hydrogen bromide	HBr
hydrogen iodide	HI
hydrogen sulphide	H_2S

When these gases are dissolved in water an acid is formed. The acid has the same formula as the gas, but the name is changed to indicate that it is now dissolved in water.

Gas Formula	*Acid Formula*	*Acid Name*
HF	HF	hydrofluoric acid
HCl	HCl	hydrochloric acid
HRr	HBr	hydrobromic acid
HI	HI	hydroiodic acid
H_2S	H_2S	hydrosulphuric acid

The root name of the acid is taken from the non-metal present. Remember the prefix hydro means *binary acid*.

Ternary Acids

Ternary acids contain three elements, one of which is always hydrogen. For the moment we will confine ourselves to those ternary acids that also contain oxygen. These are the *oxy-acids*. The third element present gives the acid its root name.

For example:

Acid	*Formula*
fluoric acid	HFO_3
chloric acid	$HClO_3$
bromic acid	$HBrO_3$
iodic acid	HIO_3
nitric acid	HNO_3
sulphuric acid	H_2SO_4
carbonic acid	H_2CO_3
phosphoric acid	H_3PO_4

Each of these acids is a member of a different family of acids. The families are similar in that they all contain the same three elements, but they differ in the number of oxygen atoms contained in the molecule. Remember the example of sulphur a while back?

Here is the sulphuric acid family:

H_2SO_5 — persulphuric acid
H_2SO_4 — sulphuric acid
H_2SO_3 — sulphurous acid
H_2SO_2 — hyposulphurous acid

In each oxy-acid the central non-metallic element provides the root name. The prefix and suffix to the root give the number of oxygen atoms as compared to the parent acid.

If the parent acid is phosphoric acid, H_3PO_4, then phosphorous acid has one less oxygen, H_3PO_3, perphosphoric acid has one more oxygen than the parent acid, H_3PO_5. Hypophosphorous acid has two fewer oxygens than the parent and therefore is H_3PO_2.

All you really need to know are the formulae for the parent acids and these simple rules for naming.

Problems

1. Write the name and formula for all members of the carbonic acid family.
2. If telluric acid has the formula H_2TeO_4, write the formulae and names of other members of its family.
3. Give the names and formulae of the substances formed when:
   (a) magnesium oxide is dissolved in water;
   (b) carbon dioxide is dissolved in water;
   (c) barium oxide is dissolved in water;
   (d) sulphur (VI) oxide is dissolved in water;
   (e) sodium is burned in air.
4. What acid might you expect to be present in your blood stream?
5. What acid is present in your stomach?
6. If permanganic acid is $HMnO_5$, then give the names and formulae of other members of the family.

Acid Strength and Base Strength

In determining the properties of acids in Experiment 4.4 you observed that all acids have many properties in common. They differ, however, in the rate at which they react. For example, when the carbonate was added to hydrochloric acid the bubbling was much more vigorous than when it was added to acetic acid. But how can this be so, if both acids had the same amount of acid initially? The answer must lie in what each acid has in common and what happens to the common part in water.

The four acids used earlier were:

hydrochloric	HCl
sulphuric	H_2SO_4
nitric	HNO_3
acetic	CH_3COOH

You will notice that each of these acids contains hydrogen. Since each has many properties in common it would seem that the hydrogen part of each acid must be responsible for these common properties. But how do we explain differences in reaction rate for acids of the same concentration?

Recall at this time Experiment 3.9 on electrical conductivity and the theory that was proposed to explain the conductivity of electricity in solutions. With this in mind try the following experiment.

4.7 Electrical Conductivity of Acids and Bases

conductivity testing apparatus
hydrochloric acid (0.1 M, 1 M and 5 M)
acetic acid (0.1 M, 1 M and 5 M)
sulphuric acid (0.1 M)
ammonium hydroxide (0.1 M)
potassium hydroxide (0.1 M)
calcium hydroxide (saturated)
sodium hydroxide (0.1 M)
barium hydroxide (saturated)
distilled water

1. Use the same electrical conductivity apparatus used in Experiment 3.9 and test the conductivity of each of the solutions listed above. Remember to rinse the electrodes in distilled water after testing the conductivity of each solution. List the solutions under the headings of *Good, Fair* or *Poor* conductors of electricity.

Questions

1. Why do the solutions conduct an electrc current? (Refer to your earlier conductivity experiment for the answer.)
2. On the basis of your answer to Question 1 account for the fact that the 1 M HCl solution is a better conductor than the 0.1 M HCl.
3. On the basis of your answer to Question 2 account for the fact that the 1 M HCl solution is a better conductor than the 1 M CH_3COOH.

Acid Strength

A brief statement from Arrhenius' Theory of Ionization might help to clarify the results of this experiment.

"Some molecules dissociate into charged particles called ions."

Solutions that do not conduct electricity have no ions present. Solutions that do conduct electricty have ions present. Would it, therefore, not follow that solutions that conduct very well have more ions present than those that do not conduct well?

The simplest assumption would seem to be that solutions that conduct well contain many ions. Solutions that are only fair conductors have fewer ions.

But how can we explain that 0.1 M hydrochloric acid conducts so much better than 0.1 M acetic acid?

There must be more ions present in 0.1 M hydrochloric acid than 0.1 M acetic acid. Since both solutions contain the same number of molecules, it therefore follows that fewer acetic acid molecules dissociate into ions. Put another way, only some acetic acid molecules ionize. Some remain in the solution as molecules.

An acid that is highly ionized is called a *strong acid*.

A base that is highly ionized is call a *strong base*.

An acid or base that is only slightly ionized is said to be *weak*.

We should consider one more thing. 0.1 M hydrochloric acid reacts more rapidly than 0.1 M acetic acid. As we have just seen, hydrochloric acid is much more ionized than acetic acid. It seems reasonable, then, to assume that the reactions characteristic of an acid are reactions involving the ions produced by the acid. They are not reactions that involve the un-ionized molecule.

Since all acids have the same reactions they must all produce the same ion or ions. The only ion that all acids have in common is the hydrogen ion, H^+. We generally consider an acid to be any molecule that can produce hydrogen ions. The more hydrogen ions that are produced, the stronger the acid.

Without trying to complicate the issue we must add one more detail. When a molecule dissolves in water to produce a hydrogen ion, the hydrogen ion immediately joins up with a water molecule.

$H^+ + H_2O \longrightarrow H_3O^+$

This new ion is called a *hydronium* ion. It is a *hydrated hydrogen ion*. Most hydrogen ions in a solution are hydrated.

The same sort of arguments apply to bases. In this case, however, the ion responsible is the hydroxyl ion, OH^-. Hydroxyl ions are not generally thought to hydrate.

Since all acidic solutions contain the hydrogen ion (hydronium ion) and all basic solutions contain the hydroxyl ion, let's see what would happen if we mixed these two solutions.

Neutralization

4.8

Neutralization Reaction between an Acid and a Base

evaporating dish
dropper
retort stand and ring
wire gauze

bunsen burner
1 M sodium hydroxide
1 M hydrochloric acid
phenolphthalein solution

1. Add 5 ml of 1 M sodium hydroxide to a clean evaporating dish. Add three drops of colourless phenolphthalein indicator to the base in the evaporating dish.
2. Add 1 M hydrocholoric acid solution to the base drop by drop. Stir the solution all the time until the colour just disappears.
3. Place the evaporating dish and solution in it on a wire gauze that is resting on a retort ring . Heat the evaporating dish until the solution has evaporated to dryness.

As the dish nears dryness there may be some spattering. Heat the dish very gently at this stage in the evaporation.

4. Describe the residue. Taste a small portion of it.

Warning

Be careful — the dish will be hot!

Questions

1. What is the function of the phenolphthalein in the reaction?
2. Was the solution acidic or basic after the acid had been added? Explain.
3. Can you name any common substance that tastes like the residue?
4. Write a balanced chemical equation for the reaction that took place between the HCl and the NaOH.

Tests on the product of this reaction show that it is neither acidic nor basic. Such a solution is said to be *neutral*. The process is called *neutralization*. What has happened to the original solutions?

We began with hydrogen ions and chloride ions from the acid.

$HCl \longrightarrow H^+ + Cl^-$

Notice that we are not showing the H^+ as hydrated.

From the base came sodium ions and hydroxyl ions.

$NaOH \longrightarrow Na^+ + OH^-$

You were able to identify one of the products as salt — sodium chloride. That can only mean that the hydrogen ion and the hydroxyl ion combined.

$H^+ + OH^- \longrightarrow H_2O$

Since all acids contain hydrogen ions and all bases contain hydroxyl ions, all neutralization reactions must produce water. There will always be a second product as well. This other product will contain the positive ion from the base and the negative ion from the acid. This other product is called a *salt*.

Neutralization can then be defined as the reaction between an acid and a base to produce a salt and water.

Questions

Write balanced formula equations for the following reactions. Name the salt produced in each case.

1. Potassium hydroxide reacts with hydrochloric acid.
2. Lithium hydroxide reacts with hydrochloric acid.
3. Calcium hydroxide reacts with hydrosulphuric acid.
4. Barium hydroxide reacts with hydroiodic acid.
5. Aluminum hydroxide reacts with hydrobromic acid.

From these equations we see that there is a 1:1 relationship between the hydrogen ion and the hydroxyl ion. If we neutralize a substance with a solution of known concentration we might be able to determine the concentration of the original solution.

When a neutralization is carried out in a *quantitative* manner it is called a *titration*. In order to achieve this neutralization we rely on the fact that a neutral solution is neither acidic nor basic. At this neutral point there must be the same number of hydrogen ions and hydroxyl ions.

We can tell we have reached this point when an indicator changes its colour. This point of transition is called the *end point*. Enough of this for now.

We need two things to perform a titration: solutions of known concentration and a way of accurately measuring liquid volumes. First the solutions.

Molarity (M) and Molar Solutions

Many of the solutions that you have used in your course so far have been labelled with a number followed by a capital M. The letter symbolizes molarity which is defined as moles of solute per litre of solution. For example. 1 M NaOH solution would contain 1 mol of NaOH in every litre of solution (1 mol NaOH/ℓ). To get a better understanding of molarity try the following experiment.

In the laboratory many reactions are studied by mixing solutions of the reactants. The volumes of the solutions can be easily and accurately measured. The amount of reactant used is directly proportional to the volume of the solution used. The exact amount of reactant used may be calculated from the volume of solution used provided you know the concentration of the solution.

In this experiment you will prepare aqueous solutions of sodium hydroxide and hydrochloric acid.

4.9 Preparing a Molar Solution

balance
2 100 ml volumetric flasks
graduate cylinder
sodium hydroxide pellets
concentrated hydrochloric acid
distilled water

Part A

1. Calculate the mass of 0.1 mol of NaOH (0.1 × g in 1 mol).
2. Take this amount of NaOH and dissolve it in about 80 ml of distilled water in a 100 ml volumetric flask. When the NaOH has dissolved add distilled water to increase the volume to 100 ml.

Warning

Sodium hydroxide is extremely caustic and must not be handled with your hands.

3. The solution that you have prepared has 0.1 mol NaOH/100 ml of solution. Calculate the number of moles of NaOH that would be in 1 ℓ of the solution.
4. The number of moles in 1 ℓ of the solution is called the molarity of the solution. Label the solution as to solute and its molarity.

Part B

1. Concentrated hydrochloric acid contains 12 mol HCl/ℓ of solution. Calculate the volume of concentrated HCl solution that would contain 0.1 mol of HCl.

Volume of concentrated HCl containing 0.1 mol of HCl = 0.1 mol × 1000 ml/12 mol.

2. Carefully measure out this volume of concentrated hydrochloric acid. Use your graduated cylinder.

Warning

Concentrated acids may cause severe burns or damage to clothes and floors. Extreme care must be taken to avoid spillage.

3. Add the acid from Step 2 to a 100 ml volumetric flask. Add distilled water to bring the volume to 100 ml.

4. The acid solution that you have prepared contains 0.1 mol HCl in 100 ml of solution. Calculate the number of moles that would be in 1 ℓ of the solution. Label this solution as to solute and molarity.
5. Keep both solutions for your next experiments.

It is very important that you understand that even though the solutions that you have prepared each contain only 0.1 mol of solute each is labelled as a 1 M solution. The labelling of the solution refers to the moles of solute that would be in a litre of the solution whereas the former refers to the number of moles of solute in the volume of solution that you have prepared.

Molar Solutions

Example 1

Explain how you would prepare 1 ℓ of 5 M potassium hydroxide solution.

5 M means there are 5 mol in every litre of solution (5 mol/ℓ). We are required to prepare 1 ℓ. Therefore we will need 5 mol of potassium hydroxide.

Potassium hydroxide has the formula KOH. The gram molecular mass will be the sum of the gram atomic masses.

$$\overset{\text{K}}{39.1\ \text{g}} + \overset{\text{O}}{16.0\ \text{g}} + \overset{\text{H}}{1.0\ \text{g}} = 56.1\ \text{g}$$

1 mol of KOH has a mass of 56.1 g.

5 mol of KOH has a mass of 56.1 g × 5 = 280.5 g

To prepare the solution dissolve 280.5 g of solid potassium hydroxide in less than 1 ℓ of distilled water. After the solid has dissolved add enough distilled water to make exactly 1 ℓ of solution. Label the solution 5 M KOH.

Example 2

Explain how you would prepare 5 ℓ of 0.5 M sodium iodide solution.

0.5 M means 0.5 mol/ℓ. Therefore in 5 ℓ we will require

$$\frac{0.5\ \text{mol}}{1\ \ell} \times 5\ \ell = 2.5\ \text{mol}$$

One mole of sodium iodide has a mass of

$$\overset{\text{Na}}{23.0\ \text{g}} + \overset{\text{I}}{126.9\ \text{g}} = 149.9\ \text{g}$$

Therefore 2.5 mol will have a mass of

$$\frac{149.9\ \text{g}}{1\ \text{mol}} \times 2.5\ \text{mol} = 374.8\ \text{g}$$

To prepare the solution dissolve 374.8 g of sodium iodide in less than 5 ℓ of distilled water. When the solute has dissolved add enough distilled water to make exactly 5 ℓ of solution. Label this solution 0.5 M NaI.

Example 3

Explain how you would prepare 100 ml of 0.1 M tin (II) fluoride solution.

0.1 M means 0.1 mol/ℓ. Therefore in 0.1 ℓ (100 ml) we will require

$$\frac{1}{10}\,\ell \times \frac{0.1\ \text{mol}}{\ell} = 0.01\ \text{mol}$$

One mole of tin (II) fluoride has a mass of

$$\overset{\text{Sn}}{118.7\ \text{g}} + \overset{\text{F}_2}{2(19.0)\ \text{g}} = 156.7\ \text{g}$$

Therefore 0.01 mol will have a mass of

$$\frac{156.7\ \text{g}}{1\ \text{mol}} \times 0.01\ \text{mol} = 1.57\ \text{g}$$

To prepare the solution dissolve 1.57 g tin (II) fluoride in about 90 ml of distilled water, then add enough water to make exactly 100 ml. Label the solution 0.1 M SnF_2.

Problems

1. Give instructions how to prepare the following solutions:
   (a) 1.0 ℓ of 1.0 M NaCl
   (b) 3.0 ℓ of 2.5 M $AgNO_3$
   (c) 1.5 ℓ of 0.35 M $Pb(ClO_3)_2$
   (d) 250 ml of 0.40 M $Ca(OH)_2$
   (e) 100 ml of 0.002 M $Fe(SCN)_2$
2. Hydrochloric acid is purchased as a liquid 12 M solution. How would you prepare 1 ℓ of 1 M HCl solution?
3. Sulphuric acid is purchased as a liquid 18 M solution. How would you prepare 1 ℓ of 1 M H_2SO_4 solution?
4. Acetic acid is purchased as a liquid 17 M solution. How would you prepare 0.5 ℓ of 5 M CH_3COOH?
5. Nitric acid comes as a 16 M stock solution from the manufacturer and is normally sold in 3 kg containers. The specific gravity of 16 M nitric acid is 1.42. How would you prepare 1 ℓ of 2 M HNO_3?
6. Liquid ammonium hydroxide (NH_4OH) is sold as a solution that is 58 percent NH_4OH. The specific gravity of the solution is 0.90. How would you prepare 1 ℓ of 1 M NH_4OH solution?

Table 4.3

Reagent	*Formula*	*Stock Molarity*	*Specific Gravity*	*Percentage Solute*
Acetic acid	CH_3COOH	17 M	1.05	99%
Hydrochloric acid	HCl	12 M	1.20	35%
Nitric acid	HNO_3	16 M	1.40	70%
Sulphuric acid	H_2SO_4	18 M	1.85	96%
Ammonium hydroxide	NH_4OH	15 M	0.90	58%

We have decided to express our concentrations using molarity. It seems to be most convenient for us at the moment. Please remember that solution concentrations can be expressed in three different ways:

1. molarity;
2. percentage;
3. $\frac{\text{grams of solute}}{\text{100 g of solvent}}$ (solubility units)

Having mastered the concept of molarity and having prepared your own molar solutions of hydrochloric acid and sodium hydroxide it is time to look at a titration.

If you have not already done so, turn to Skills, Unit 8, and read up on liquid volume measurement. Look especially at the pipette and the burette.

Titrations

4.10

Acid-Base Titration

retort stand
burette
burette clamp
pipette and bulb
125 ml erlenmeyer flask
distilled water
acid and base solutions from Experiment 4.9
phenolphthalein solution

1. Rinse all the equipment that you will be using with distilled water. This will include a pipette, a burette and a 125 ml erlenmeyer flask.
2. Mount your burette in a clamp as shown in the diagram (stopcock closed). Fill your burette with the acid solution from Experiment 4.9. Pour acid solution into the burette until the liquid is about 2 cm from the opened end of the burette. Open the stopcock carefully and allow acid solution to run into the beaker until the liquid in the burette has moved down to the zero reading on the burette. Read from the bottom of the meniscus (see diagram).
3. Using a pipette and bulb carefully measure out 10 ml of base solution and put it in a clean 125 ml erlenmeyer flask.
4. Add three drops of the phenolphthalein to the flask.
5. Put the mouth of the flask under the tip of the burette and add acid from the burette a drop at a time. Continually swirl the flask while the acid is being added. Continue to add the acid until the colour of the indicator just disappears.

6. Record the volume of acid required just to neutralize 10 ml of the base solution.
7. Repeat the same procedure twice more. It is not necessary to refill the burette. Just record the starting volume before each of the remaining titrations.
8. Save the remaining acid for the next experiment.

Questions

1. Why is it necessary to fill the tip of the burette before beginning a titration?
2. What advantages are there to using a pipette bulb rather than oral suction to fill the pipette?
3. Calculate the number of moles of NaOH in 10 ml of the 1 M NaOH solution.

Moles of solute in 10 ml = molarity × 10 ml volume/1000.

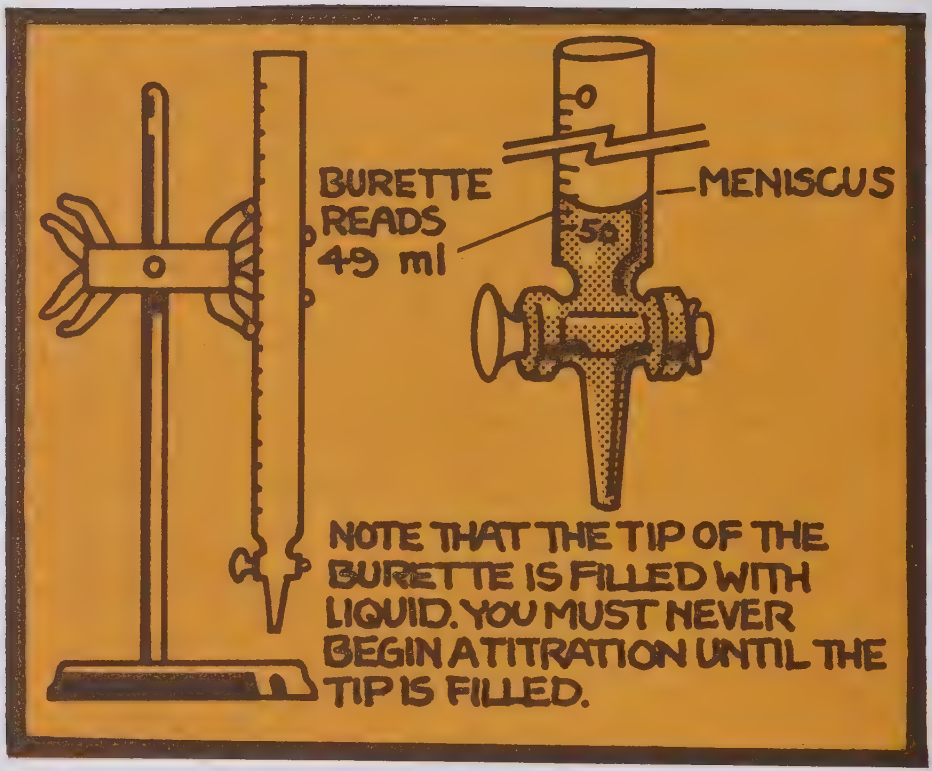
4. Calculate the number of moles of HCl in 10 ml of the 1 M HCl solution (use average volume of HCl solution).
5. Write the equation for the neutralization reaction.
6. Calculate the ratio: moles HCl used/moles NaOH used.
7. From the equation determine the ratio in moles in which HCl reacts with NaOH.
8. How does the ratio from Question 6 compare to the ratio from Question 7?
9. If the HCl solution used in the burette had been a 0.1 M solution what volume of solution would have been used?
10. If the NaOH solution had been a 0.1 M solution what volume of 1 M acid solution would have been used?

4.11

Determination of the Molarity of a Potassium Hydroxide Solution from an Acid of Known Molarity

apparatus as in Experiment 4.10
0.1 M HCl solution
KOH solution of unknown molarity
phenolphthalein solution

1. Rinse all glassware with distilled water.
2. Fill your burette with the HCl solutions as in Experiment 4.10.
3. Measure out 10 ml of KOH solution with a pipette. Add three drops of phenolphthalein.


BURETTE READS 49 ml
MENISCUS
0
50
NOTE THAT THE TIP OF THE BURETTE IS FILLED WITH LIQUID. YOU MUST NEVER BEGIN A TITRATION UNTIL THE TIP IS FILLED.

4. Add acid from the burette until the colour of the indicator just disappears. Record the volume of the acid used.
5. Repeat the procedure twice more.

Questions

1. Determine the average value for the volume of acid used.
2. Calculate the moles of HCl in the volume of acid used. (Refer to the questions of Experiment 4.10.)
3. Calculate the moles of H^+ ion used in the reaction.
4. Calculate the moles of OH^- ion that must have been in 10 ml of the KOH solution.
5. Calculate the moles of KOH that must have been in 10 ml of the KOH solution.
6. Calculate the moles of KOH that would be in 1 ℓ of the KOH solution. This is the molarity of the base solution.

Lab Test

Your teacher will provide you with a different unknown and a standardized solution. Determine the concentration of this unknown. Hand in your data, calculation and conclusion before the end of the period.

The acid solution used in Experiment 4.11 is called a *standard solution*. A standard solution is a solution of known concentration. The standard solution is usually the titrating reagent (the solution in the burette), while the substance being determined is the titrated reagent. In Experiment 4.11 the KOH solution is the titrated reagent.

Example 1

1. How many moles of hydrogen ion would be available from 100 ml of 0.1 M HCl solution:

   0.1 M means 0.1 mol/ℓ of solution. Therefore 100 ml (0.1 ℓ) contains:

   $$\frac{1\ \ell}{10} \times \frac{0.1\ \text{mol}}{\ell} = 0.01\ \text{mol HCl}$$

   One mole of HCl contains 1 mol of H^+. Therefore 0.01 mol of HCl contains 0.01 mol H^+.

   0.01 mol hydrogen ion are available.

2. How many moles of hydrogen ion would be available from 100 ml of 0.1 M H_2SO_4 solution:

   0.1 M means 0.1 mol/ℓ. Therefore 100 ml contains:

   $$\frac{1\ \ell}{10} \times 0.1\ \text{mol} = \frac{0.01\ \text{mol}\ H_2SO_4}{\ell}$$

   One mole of H_2SO_4 contains 2 mol of H^+. Therefore 0.01 mol of H_2SO_4 contains 0.02 mol H^+.

   0.02 mol hydrogen ion are available.

Example 2

A 10 ml sample of base is titrated to the end point with 25 ml of 0.1 M HCl. Calculate the concentration of the base.

First consider the acid alone. 25 ml of 0.1 M HCl contains:

$$\frac{25\ \ell}{1000} \times \frac{0.1\text{mol}}{\ell} = 0.0025\ \text{mol HCl}$$

0.0025 mol HCl contain 0.0025 mol H^+ at the end point.

$$\text{total moles } H^+ = \text{total moles } OH^-.$$

Therefore the 10 ml sample of base must have contained 0.0025 mol of OH^-.

The concentration of the base must have been:

$$\frac{0.0025\ \text{mol}}{\frac{10}{1000}\ \ell} = 0.25\ \text{M in } OH^-$$

Example 3

A 10 ml sample of acid is titrated against 0.5 M NaOH. A total of 23.5 ml base are required to reach the end point. Calculate the concentration of the acid.

First consider the base alone. 23.5 ml of 0.5 M NaOH contains:

$$\frac{23.5\ \ell}{1000} \times \frac{0.5\ \text{mol}}{\ell} = 0.01175\ \text{mol NaOH}$$

0.01175 mol NaOH contain 0.01175 mol OH^-. At the end point, total moles OH^- = total moles H^+.

Therefore the 10 ml sample of acid must have contained 0.01175 mol H^+. The concentration of the acid must have been:

$$\frac{0.01175\ \text{mol}}{\frac{10}{1000}\ \ell} = 1.175\ \text{M in H}^+ = 1\ \text{M}$$

Problems

1. Calculate the molarity of an unknown base if a 10 ml sample of the base required 15 ml of 0.5 M HCl to neutralize it.
2. Calculate the molarity of an unknown acid if a 10 ml sample of the acid required 25 ml of 0.3 M NaOH to neutralize it.
3. In order to determine the concentration of citric acid in a can of grapefruit juice concentrate, three separate 10 ml samples of the concentrate were titrated against 0.1 M KOH. The amount of base required to neutralize each sample was 16.3 ml, 17.1 ml, 16.5 ml respectively. Calculate the molarity of acid in the concentrate.
4. Soap is manufactured from fats and a strong base. If too much base is left unreacted in the soap it is hard on the skin. In order to determine the amount of base present in a bar of soap the following experiment was performed.

   A 100 g bar of soap was cut into approximately 10 g sections. The mass of each section was determined and then three sections were dissolved in different 50 ml samples of distilled water. Each of these solutions was titrated against 0.01 M HCl solution. Here are the results.

Sample	*Mass*	*Volume of Acid Required to Neutralize*
1	9.3 g	5.1 ml
2	10.1 g	5.5 ml
3	12.5 g	6.8 ml

   Estimate the total number of grams of base left unreacted in the soap bar, assuming the base was NaOH.
5. A chemistry student accidentally dropped and broke a 1 ℓ container of 1 M HCl on the lab floor. Approximately how much calcium hydroxide should be spread over the spill area before any clean up is attempted?

4.12 Acidity of Some Common Substances

Earlier in our study of acids and bases we found a number of common substances that were acidic in nature. The object of our next experiment is to determine just how acidic some of these substances really are. That is, we want to find out how many moles of hydrogen ion are in a litre of some of these substances.

apparatus as in Experiment 4.10
standard base (0.1 M NaOH solution)
phenolphthalein solution
common substances provided by the student, such as vinegar, soft drink, fruit juice and ketchup

1. Follow a procedure similar to that of Experiment 4.10. Since the NaOH solution is the standard this time it will be placed in your burette.
2. Add 5 ml of the common substance to the 125 ml erlenmeyer flask with three drops of the indicator. Add base until a colour just appears. Record the volume of standard base used.
3. Repeat with other substances as time permits.

Questions

1. Calculate the number of moles of NaOH used.
2. Calculate the number of moles of OH^- ion reacting.
3. Calculate the number of moles of H^+ ion that must have reacted.
4. Calculate the number of moles of H^+ ion that must be in 1 ℓ of the common substance.
5. Use the class results to compare the relative acidity of the common substances used.
6. What major problems did you encounter?

4.13 Back Titrations

A different aspect of neutralization titrations is what is called a back titration. The concept is best illustrated by an experiment.

apparatus as in Experiment 4.10
0.1 M HCl
NaOH pellet
0.1 M NaOH standard solution
phenolphthalein solution

1. Rinse all glassware with distilled water.

2. Measure out 10 ml of the acid solution with a pipette and put it in a 125 ml erlenmeyer flask. Add one sodium hydroxide pellet to the acid in the flask. Swirl the flask until the pellet has dissolved. Add three drops of phenolphthalein indicator.

You will note that the solution is still acidic. To determine how much acid was neutralized by the pellet, titrate the solution with standard NaOH.

3. Titrate with 0.1 M NaOH until the indicator colour just appears.
4. Record the volume of standard base required to neutralize the excess acid.

Questions

1. Calculate the number of moles of HCl in the 10 ml sample from Step 2.
2. Calculate the number of moles of NaOH in the volume of 0.1 M NaOH required to neutralize the excess acid after the pellet has dissolved.
3. Calculate the number of moles of HCl neutralized by the NaOH pellet (moles of HCl initially − moles of NaOH added from burette).
4. Calculate the number of moles of OH^- ion provided by the pellet.

In a back titration the substance being investigated is not titrated directly. In such a titration the substance under investigation is reacted with a given volume of solution of known molarity. The unreacted solution is then titrated with a second solution of known molarity in order to determine how much of the initial solution had been used.

Back titrations are used where a direct titration is not convenient. The next experiment will require back titrations.

4.14

Investigation of Several Commercial Antacid Preparations

appartus as in Experiment 4.10
1.0 M HCl
1.0 M NaOH
antacid preparations
phenolphthalein solution

1. Prepare a model "stomach" by adding 10 ml of 1.0 M HCl to an erlenmeyer flask. Add one antacid tablet and allow the reac-

tion to subside. Add three drops of phenolphthalein and back titrate the excess acid with the standard base (1.0 M NaOH). Repeat the procedure twice more for the same antacid tablet.

2. Repeat a similar procedure for other antacid tablets as instructed by your teacher.

Questions

1. Calculate the number of moles of HCl that were in the initial 10 ml sample.
2. How many moles of H^+ ion were in the initial 10 ml sample?
3. Calculate the number of moles of NaOH that were in the volume of base required for the back titration.
4. How many moles of H^+ were neutralized by the back titration?
5. How many moles of H^+ were neutralized by the antacid preparation used?
6. List the antacid tablets in order of their effectiveness at neutralizing acid.

At this point we must try to make one thing clear. One should not judge an antacid preparation on the basis of its neutralizing capacity alone. If that were the sole criteria then we would use lye for upset stomachs. Remember that besides neutralizing acid the preparation must not damage your tissues in any way. It must not upset your system any further. It must not have any unpleasant side effects. This lab study has shown you how to check just one aspect of antacid preparations. There are many more things to consider.

4.15

The Reaction of Some Metals with Dilute Acids

4 test tubes
splints

0.5 M HCl
0.5 M H_2SO_4
0.5 M CH_3COOH
calcium
zinc
iron
magnesium ribbon

1. Set up four clean test tubes and add to each 3 ml of 0.5 M hydrochloric acid. To Tube 1 add a lump of calcium; to Tube 2, a lump of zinc; to Tube 3, a piece of iron and to Tube 4, a strip of magnesium ribbon. As the reaction is progressing in each test tube bring a burning splint to the mouth of each test tube. Record any similarities and differences in each of the four test tubes.

2. Repeat Step 1 with each of the other two acids.
3. Bring a burning splint to the mouth of a test tube of hydrogen gas that is provided by your teacher.

Questions

1. Describe the gas that was produced in each test tube (colour, odour, effect on the splint).
2. Describe the hydrogen gas provided by your teacher (colour, odour, effect on the splint).
3. On the basis of the physical properties of the gas produced and the flame test, what gas is produced by the action of an acid on a metal?

Displacement Reactions

You should have no difficulty in identifying the gas produced in these reactions as hydrogen. Writing the equation may present a few problems. Let's consider the reaction between hydrochloric acid and zinc.

$HCl + Zn \longrightarrow H_2 + ??$

The unknown product should consist of the chloride ion and zinc, in which case using valences would give us:

$HCl + Zn \longrightarrow H_2 + ZnCl_2$

and balanced

$2\,HCl + Zn \longrightarrow H_2 + ZnCl_2$

This type of reaction is called a *displacement reaction*. The hydrogen in the acid is displaced by the metallic zinc. The product, zinc chloride, is a member of the salt family.

If we consider now the reaction between sulphuric acid and magnesium the equation should be:

$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$

This leads us to the method used for naming the salts of ternary acids such as $MgSO_4$. All you need to remember is the names of the parent acids and how to derive other members of the acid families.

Each ternary acid consists of two parts: the hydrogen and the other ion. The other ion is composed of several atoms and is called a *complex ion* or, more often, a *radical*.

For example: $H_2SO_4 \longrightarrow 2H^+ + SO_4^=$ ($SO_4^=$ is a *radical*)
$HClO_3 \longrightarrow H^+ + ClO_3^-$ (ClO_3^- is a *radical*)

All the radicals that come from oxy-acids take their name from the name of the acid. Look at Table 4.4.

Table 4.4

Acid	*Formula*	*Radical*	*Radical Name*
perchloric acid	$HClO_4$	ClO_4^-	perchlorate
chloric acid	$HClO_3$	ClO_3^-	chlorate
chlorous acid	$HClO_2$	ClO_2^-	chlorite
hypochlorous acid	$HClO$	ClO^-	hypochlorite

Acid radicals are named by taking the name of the acid that produces them and changing the ending. The *ic* acid ending is changed to *ate*. The *ous* acid ending is changed to *ite*. Look at Table 4.5

Table 4.5

Acid	*Formula*	*Radical*	*Radical Name*
persulphuric	H_2SO_5	SO_5^{-2}	persulphate
sulphuric	H_2SO_4	SO_4^{-2}	sulphate
sulphurous	H_2SO_3	SO_3^{-2}	sulphite
hyposulphurous	H_2SO_2	SO_2^{-2}	hyposulphite

Ternary salts are simple to name. Name the positive ion, then name the negative ion. To name the radical of an oxy-acid, change the -ic acid ending to -ate or the -ous acid ending to -ite.

$CuSO_4$	copper (II) sulphate
$MgSO_4$	magnesium sulphate
$NaClO$	sodium hypochlorite
K_2CO_3	potassium carbonate
$AlPO_5$	aluminum perphosphate

Careful examination of these examples will show that the valence number of a radical is the same as the number of hydrogen ions in the parent acid.

Questions

Name these ternary salts:

1. $CuSO_5$
2. $MgSO_3$
3. $NaClO_3$
4. K_2CO_2
5. $AlPO_4$
6. $BaSO_4$
7. $Zn(BrO_3)_2$
8. Li_2SO_5
9. $Ca(NO_3)_2$
10. $Mn(BrO_3)_2$
11. $Fe(IO_2)_3$
12. $Ni(FO_3)_2$
13. $Cu(NO_4)_2$
14. $Zn_3(PO_4)_2$

15. $AgNO_3$
16. $Pb(SO_4)_2$
17. $CrCO_3$
18. $Sr(ClO_2)_2$
19. Cs_2SO_3
20. $HgCO_4$

We might as well remember acetic acid:

$CH_3COOH \longrightarrow H^+ + CH_3COO^-$—acetate ion

Which could produce:

$Pb(CH_3COO)_2$	lead (II) acetate
$AgCH_3COO$	silver acetate
$NaCH_3COO$	sodium acetate

And don't forget ammonium hydroxide which gives us two other radicals:

$NH_4OH \longrightarrow NH_4^+ + OH^-$

ammonium hydroxide

which could produce:

NH_4NO_3	ammonium nitrate
$(NH_4)_3PO_4$	ammonium phosphate
$CsOH$	cesium hydroxide

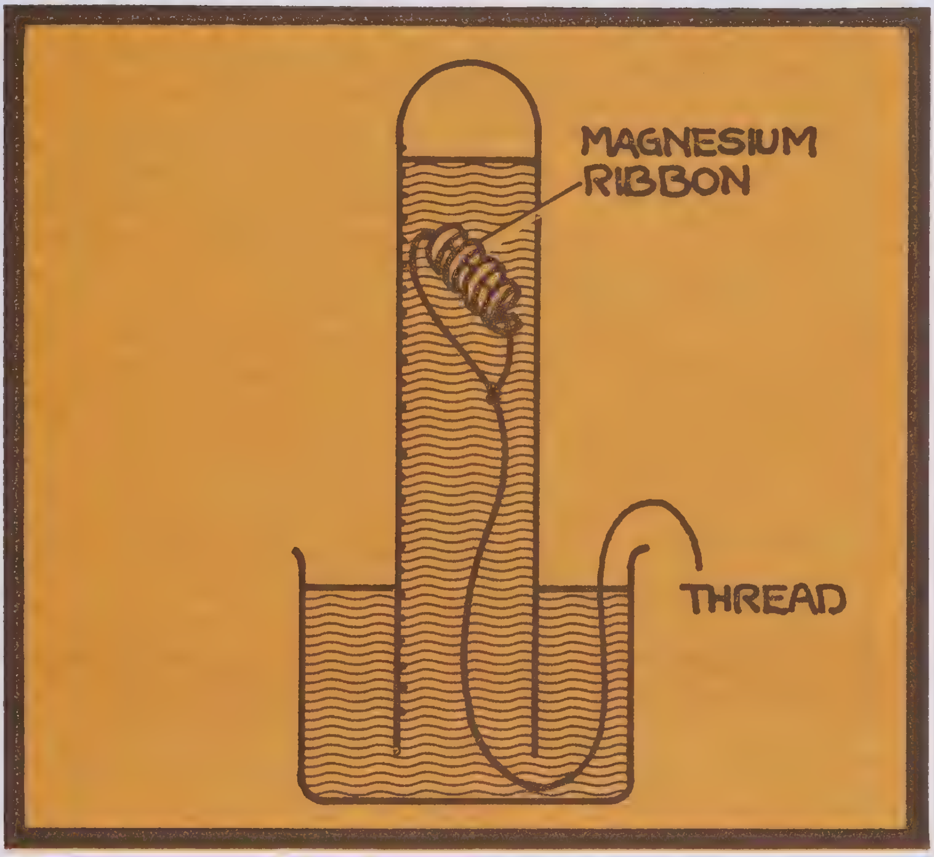
At this point in the unit you should know and be able to do the following:

1. You should know the names and the formulae for the parent oxy-acids.
2. Be able to derive the names and formulae for other members of any acid family.
3. Be able to derive the name, formula and valence number of any radical produced by an oxy-acid.
4. Be able to name the salts produced by the oxy-acids in neutralization or displacement reactions.
5. Be able to write balanced chemical equations to represent neutralization or displacement reactions.

4.16 Investigating Equations

We have looked qualitatively at the reactions between metals and acids. Its now time to gather some quantitative information.

magnesium ribbon
500 ml beaker
burette, gas measuring cylinder or graduated cylinder
5 M hydrochloric acid
thermometer
thread — 50 cm long



1. Cut a 5 cm length of magnesium ribbon. Your teacher will give you the mass of a 100 cm length of ribbon.
2. Wind the magnesium into a spiral around the end of a pencil or glass rod. Tie the thread through the spiral and knot it securely.
3. Pour 20 ml of 5 M HCl into the bottom of the burette (or gas measuring tube or graduated cylinder) carefully and slowly so as to avoid mixing. Fill the rest of the container with distilled water.
4. Invert the burette into a beaker three-quarters filled with water. Avoid any air bubbles if you can. If not, record the volume of trapped air.
5. Gently push the magnesium coil into the open end of the inverted burette. The reaction will not begin immediately but when it does use the thread to prevent the coil from sticking to the walls of the burette.
6. When the reaction has stopped and all the magnesium has reacted record:
   (a) the volume of trapped gas;
   (b) the temperature of the water;
   (c) atmospheric pressure;
   (d) vapour pressure of water at appropriate temperature;
   (e) difference in level between the liquid in the burette and the liquid in the beaker.

Questions

1. Determine the mass of the magnesium ribbon.
2. Determine the number of moles of magnesium used.
3. Determine the volume of hydrogen gas collected by subtracting any trapped air from the final volume.
4. Determine the pressure inside the tube by subtracting the vapour pressure of water from the barometric pressure.
5. Convert the volume of trapped hydrogen gas to STP conditions.
6. Calculate the number of moles of hydrogen produced in the reaction.
7. Compare the moles of hydrogen produced to the moles of magnesium consumed.
8. Write a balanced formula equation for the reaction.
9. Within experimental error, how do the coefficients in the balanced equation compare with the mole ratio obtained in Question 7?

4.17

Preparation and Properties of Hydrogen

We are going now to prepare and collect hydrogen. The gas will be prepared by reacting a metal with a dilute acid. We must select a suitable metal. We want a fairly rapid reaction but not a violent reaction. Potassium would react too rapidly while lead would react too slowly. Even iron reacts slowly in a dilute acid. We will use zinc.

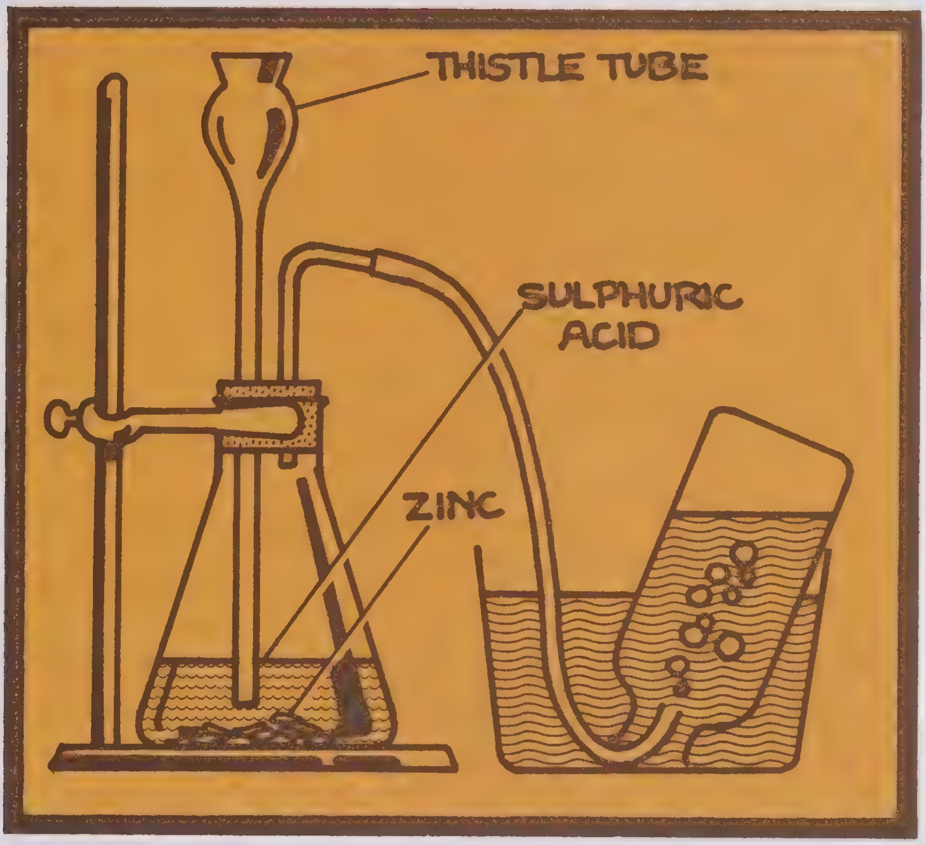
retort stand and clamp
250 ml erlenmeyer flask
thistle tube
delivery tube
stopper
pneumatic trough
3 gas bottles

3 glass plates
large beaker
splints

technical grade zinc
4 M H_2SO_4

Warning

Do not bring an open flame near the gas generator while it is in operation.



1. Set up a gas generator as illustrated in the diagram. Put five or six lumps of technical grade zinc into the flask before inserting the stopper. The end of the thistle tube should be about 0.5 cm from the bottom of the flask.
2. Invert three gas bottles filled with water into a tray of water (see diagram).
3. Carefully pour the acid solution into the flask by the thistle tube until the end of the thistle tube is immersed in the liquid. Allow the reaction to proceed for about 1 min then place the end of the delivery tube under the mouth of a gas bottle to fill it with gas. Repeat with the remaining two bottles.
4. Cover the mouth of each gas bottle and turn them upright on the desk.
5. Remove the stopper from the flask and discard the acid solution. Rinse well with water. It will now be safe to proceed with your experiment.
6. Examine the gas in one of the gas bottles. Note its colour and odour. Insert a burning splint into the mouth of this bottle. Is the gas hydrogen?
7. Invert a second bottle of gas into a beaker that is half filled with water. Swirl the bottle to mix the gas and the water.
8. Hold the third jar of gas in an inverted position and slowly insert a burning splint well into the jar; then withdraw the splint. Watch carefully.

Questions

1. Why must the end of the thistle tube be immersed in the acid solution?
2. What property of hydrogen makes it dangerous to bring an open flame near the gas generator?
3. List the following properties of hydrogen gas as determined by this experiment:
   (a) colour
   (b) odour
   (c) solubility in water
   (d) does it burn?
   (e) does it support burning?
4. Zinc metal reacts with sulphuric acid to produce zinc sulphate and hydrogen. Write the balanced chemical equation for this reaction.
5. How many moles of zinc would be required to produce 4 mol of hydrogen gas?
6. What mass of hydrogen could be produced by this reaction if we started with 25 g of sulphuric acid?
7. What volume of hydrogen at STP would be produced if we used 10 g of zinc in this reaction?

Problems

Let's close this section on names and formulae with this exercise. Write the balanced formula equation to represent the following reactions. Name *all* the products of each reaction.

1. Sulphuric acid is used to neutralize calcium hydroxide.
2. Zinc metal is dropped into a container of acetic acid.
3. Aluminum hydroxide is poured into phosphorous acid.
4. Magnesium ribbon is slowly added to a container of hydrochloric acid.
5. (a) Sodium oxide is dissolved in water.
   (b) Sulphur (IV) oxide is dissolved in water.
   (c) These two solutions are mixed together.

4.18

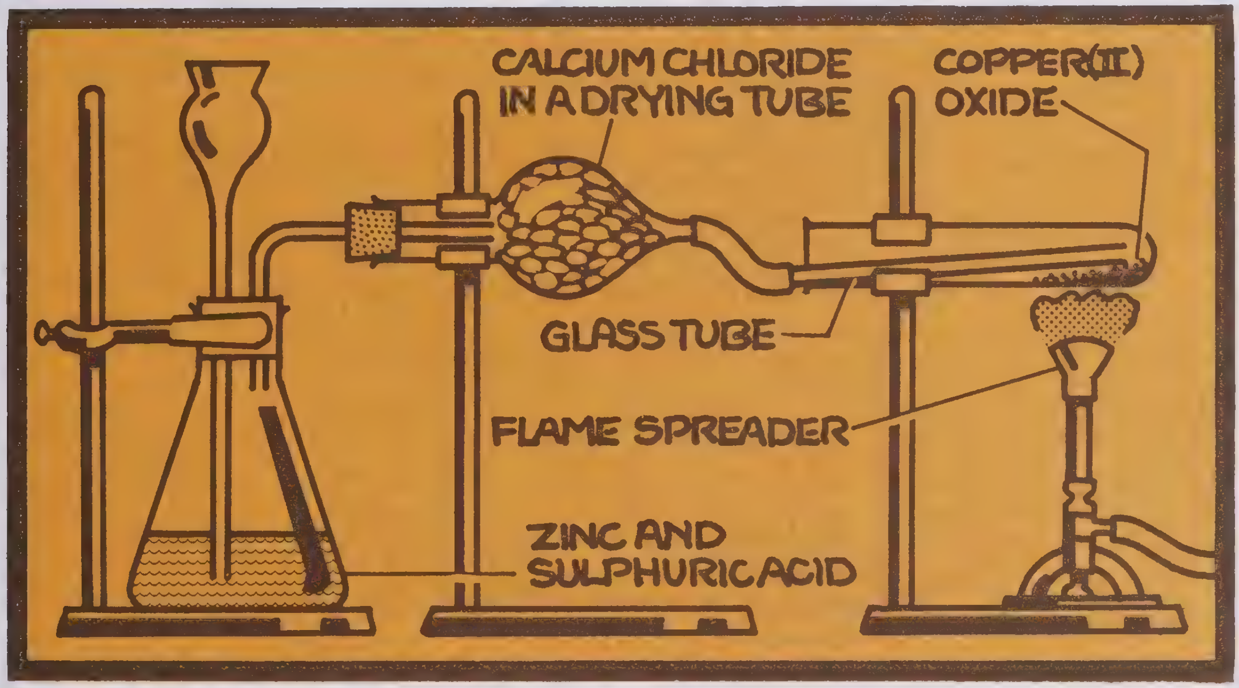
Reaction between Hydrogen and Copper (II) Oxide (Demonstration)

Earlier in this course we mixed hydrogen with oxygen in a eudiometer and ignited the mixture with a spark from an induction coil. The mixture ignited explosively producing water. On numerous occasions in the last few experiments we have seen similar examples of the strong affininty that hydrogen has for oxygen. This affinity is so great that hydrogen may even remove the oxygen from certain compounds. The reaction of hydrogen with copper (II) oxide will illustrate this property of hydrogen.

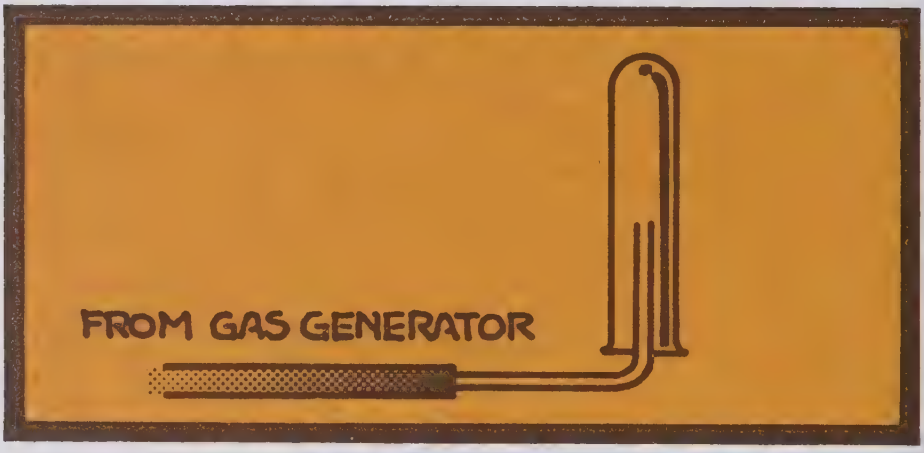
Warning

Safety glasses must be worn for this experiment.

3 retort stands with clamps
250 ml erlenmeyer flask
thistle tube
delivery tubing
stoppers
drying tube
2 test tubes
bunsen burner
flame spreader
watch glass
technical grade zinc
4 M H_2SO_4
calcium chloride
copper (II) oxide
cobalt chloride paper
concentrated nitric acid
copper shot



1. Assemble a gas generator similar to that in Experiment 4.17. Direct the delivery tube to a drying tube that is filled with calcium chloride (see diagram). Pour acid into the flask by the thistle tube until the end of the thistle tube is immersed in the liquid and a brisk effervescence results.
2. Allow gas to flow for about 2 min. Collect a test tube of gas by the downward displacement of air (see diagram). With the test tube inverted move it to a bunsen flame that is some distance from the generator. If the gas burns quietly, proceed. If it ignites explosively wait 1 min then repeat the test. Do not proceed until pure hydrogen is being delivered. Pure hydrogen will ignite quietly.



3. Allow the hydrogen to flow into the test tube with the copper (II) oxide as shown in the diagram below. At the same time begin heating the copper (II) oxide. Watch carefully and record all observations.
4. After several minutes turn off the burner and allow the test tube to cool while hydrogen is passing through it. Bring a small strip of cobalt chloride paper into contact with the liquid that is inside the test tube. Bring a similar strip of cobalt chloride paper in contact with a drop of water.
5. Remove the solid material from the test tube and place it in a watch glass. Examine it carefully. Note in particular its colour. Place a few drops of concentrated nitric acid on the residue in the watch glass. Record all observations.
6. Clean and dry the watch glass. Place a couple of pieces of copper shot in the watch glass and add a few drops of concentrated nitric acid. Record all observations.

Questions

1. In what way(s) did the residue after the reaction differ from the copper (II) oxide?
2. What happened to the cobalt chloride paper when placed in:
   (a) the liquid in the test tube?
   (b) the drop of water?
   What conclusion might we draw from this?
3. What happened when nitric acid was added to:
   (a) the residue?
   (b) the copper shot?
   What conclusion might we draw from this?
4. In your own words state what has happened to the copper (II) oxide in this experiment. What was responsible for this change?
5. Write a balanced chemical equation for the reaction between hydrogen and copper (II) oxide.
6. What mass of hydrogen would be required to react completely with 10 g of copper (II) oxide?

7. What volume of hydrogen (STP) would be required to react completely with 20 g of copper (II) oxide?

Oxidation and Reduction

Written below are chemical equations for some of the reactions that were encountered in the last 18 experiments.

$2Mg + O_2 \longrightarrow 2MgO$
$S + O_2 \longrightarrow SO_2$
$C + O_2 \longrightarrow CO_2$
$4Na + O_2 \longrightarrow 2Na_2O$
$2H_2 + O_2 \longrightarrow 2H_2O$

You will of course notice that all these reactions involve the chemical union of an element with oxygen. Such reactions are called *oxidation* reactions. The oxygen in each case is the *oxidizing agent* and the substance produced is the product of the oxidation.

An Oxidation

$2H_2$	+	O_2	$\longrightarrow$	$2H_2O$
substance oxidized		oxidizing agent		product of the oxidation

The implication made here is that oxygen is the only oxidizing agent. But this of course is untrue. If you look closely at the oxidations given above you will notice that the substance oxidized has lost electrons. For example, when neutral magnesium atoms unite with oxygen to form magnesium oxide each magnesium atom has given up two electrons to each oxygen atom. The term "oxidation" has had its meaning extended to include any reaction in which a substance loses electrons. Zinc is oxidized in the following reaction:

$Zn + S \longrightarrow ZnS$

Each zinc atom loses two electrons by donating them to a sulphur atom.

In the reaction between hydrogen and copper (II) oxide, oxygen is removed from the copper compound. Such a reaction is the exact opposite of oxidation and is termed *reduction*. As a matter of fact, the meaning of reduction has been extended to include any reaction in which a substance gains electrons (opposite of oxidation).

A Reduction

CuO	+	H_2	$\longrightarrow$	$Cu + H_2O$
substance reduced		reducing agent		product of reduction

You should realize now that you cannot have a reduction without an oxidation; the two go hand in hand.

Uses of Hydrogen

Hydrogen gas is used to inflate small weather balloons because of its low density. It is used in the manufacturing and refining of certain metals by reduction. Much of our tungsten is obtained in this way. A number of important chemicals are produced with hydrogen. Among these are ammonia, methanol and hydrogen chloride. Vegetable oils are converted to solids and used as cooking fats or in margarine by reacting the oils with hydrogen.

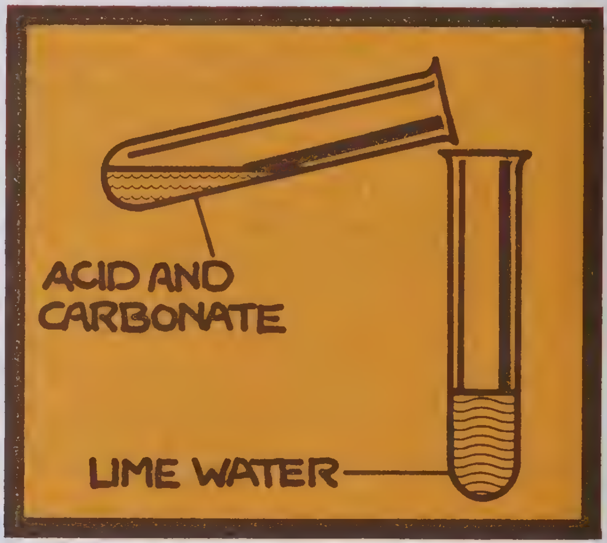
Carbon Dioxide

4.19

Reactions of Acids with Carbonates

2 test tubes
glass tube

any 3 or 4 carbonates
0.1 M HCl
0.1M H_2SO_4
0.1 M HNO_3
limewater



1. Place about 5 ml of 0.1 M HCl into an 18 × 150 mm test tube. Place about 5 ml of limewater in a second test tube. Add to the acid in the first test tube enough of one of the carbonates available to cover the tip of your spatula. As the reaction proceeds hold the two test tubes as shown in the diagram. When the reaction ceases place your thumb over the vertical test tube and shake the limewater. Record your observations.
2. Repeat Step 1 using each of the other carbonates.
3. Repeat Step 1 using each of the other acids.
4. Place 5 ml of limewater in a clean test tube and use a glass tube or drinking straw to exhale through the limewater. Record your observations.

Questions

1. What effect did your breath have on the limewater? What gas is present when you exhale?
2. What effect did the gas produced by each carbonate on each acid have on the limewater?
3. What gas is produced by the reaction of a carbonate with an acid?

The gas produced by the action of an acid on a carbonate has many important uses. A reaction similar to those observed in Experiment 4.19 is responsible for the operation of a soda-acid type fire extinguisher. The gas produced, however, is used in most fire extinguishers.

Gas Volumes from Equations

Just as we can calculate masses from an equation using the mole ratios and the molecular masses, so we can determine gas volumes from equations using the mole ratios and the molecular volume.

In the previous experiment you saw hydrochloric acid produce carbon dioxide from sodium carbonate.

$2HCl + Na_2CO_3 \longrightarrow 2NaCl + H_2O + CO_2$

The balanced equation shows the mole ratios, for example:

$$\frac{\text{moles HCl}}{\text{moles } CO_2} = \frac{2}{1}$$

By substituting molecular masses:

$$\frac{\text{grams HCl}}{\text{grams } CO_2} = \frac{73\text{ g}}{44\text{ g}}$$

However, since carbon dioxide is a gas, 1 mol occupies a volume of 22.4 ℓ at STP. If we go back to the equation and add some numbers:

$$\begin{gathered} 2HCl + Na_2CO_3 \longrightarrow 2NaCl + H_2O + CO_2 \\ 73\text{ g} + 106\text{ g} = 117\text{ g} + 18\text{ g} + 44\text{ g} \\ 179\text{ g} = 179\text{ g} \end{gathered}$$

From the information shown we could make the following statements:

(a) 2 mol of HCl produces 1 mol of CO_2
(b) 2 mol of HCl produces 44 g of CO_2

(c) 2 mol of HCl produces 22.4 ℓ of CO_2 at STP
(d) 73 g of HCl produces 1 mol of CO_2
(e) 73 g of HCl produces 44 g of CO_2
(f) 73 g of HCl produces 22.4 ℓ of CO_2 at STP

All these are equivalent statements. We can therefore write any proportion we wish that contains these relationships.

Example 1

What STP volume of carbon dioxide could be produced by 10 mol of hydrochloric acid?

2 mol of HCl produce 22.4 ℓ of CO_2 at STP
10 mol of HCl produce x ℓ of CO_2 at STP

$$\frac{2\text{ mol}}{10\text{ mol}} = \frac{22.4\ \ell}{x\ \ell}$$

$$x = 22.4\ \ell \times \frac{10}{2} = 112\ \ell$$

Therefore 10 mol of hydrochloric acid would produce 112 ℓ of carbon dioxide at STP.

Example 2

What volume of carbon dioxide would be liberated if 500 g of hydrochloric acid were poured over an excess of sodium carbonate?

73 g of HCl produces 22.4 ℓ of CO_2 at STP
500 g of HCl produces x ℓ of CO_2 at STP

$$\frac{73\text{ g}}{500\text{ g}} = \frac{22.4\ell}{x\ \ell}$$

$$x = 22.4\ \ell \times \frac{500\text{ g}}{73\text{ g}} = 153.4\ \ell = 153\ \ell$$

Therefore 500 g of hydrochloric acid would produce 153 ℓ of carbon dioxide at STP.

Example 3

What mass of hydrochloric acid would be required to produce 1000 ℓ of carbon dioxide at STP?

73 g of HCl produces 22.4 ℓ of CO_2 at STP
x g of HCl produces 1000 ℓ of CO_2 at STP

$$\frac{73\text{ g}}{x\text{ g}} = \frac{22.4\ \ell}{1000\ \ell}$$

$$x = 73\text{ g} \times \frac{1000\ \ell}{22.4\ \ell} = 3259\text{ g}$$

Therefore 3259 g of hydrochloric acid would be required to produce 1000 ℓ of carbon dioxide at STP.

Before you try the following questions remember that the molar volume *only* applies to gases.

Problems

1. What volume of sulphur dioxide gas at STP would be produced by the combustion of 1000 g of solid sulphur?
2. What mass of hydrogren peroxide would be required to produce 1 ℓ of oxygen gas at STP?
3. If 25 g of solid potassium chlorate were completely decomposed what volume of oxygen would be produced?
4. What volume of carbon dioxide would be produced if 100 g of sulphuric acid reacted completely with potassium carbonate?
5. A manufacturer wishes to make 1000 kg of solid carbon dioxide (dry ice). What volume of gaseous carbon dioxide should he begin with?
6. A student wishes to collect six 500 ml gas bottles of hydrogen gas, using zinc and 5 M sulphuric acid. Estimate the quantities of each material he should use in this exercise.
7. 100 g of hydrochloric acid are poured over 100 g of calcium carbonate. What volume of gas will be produced?

4.20

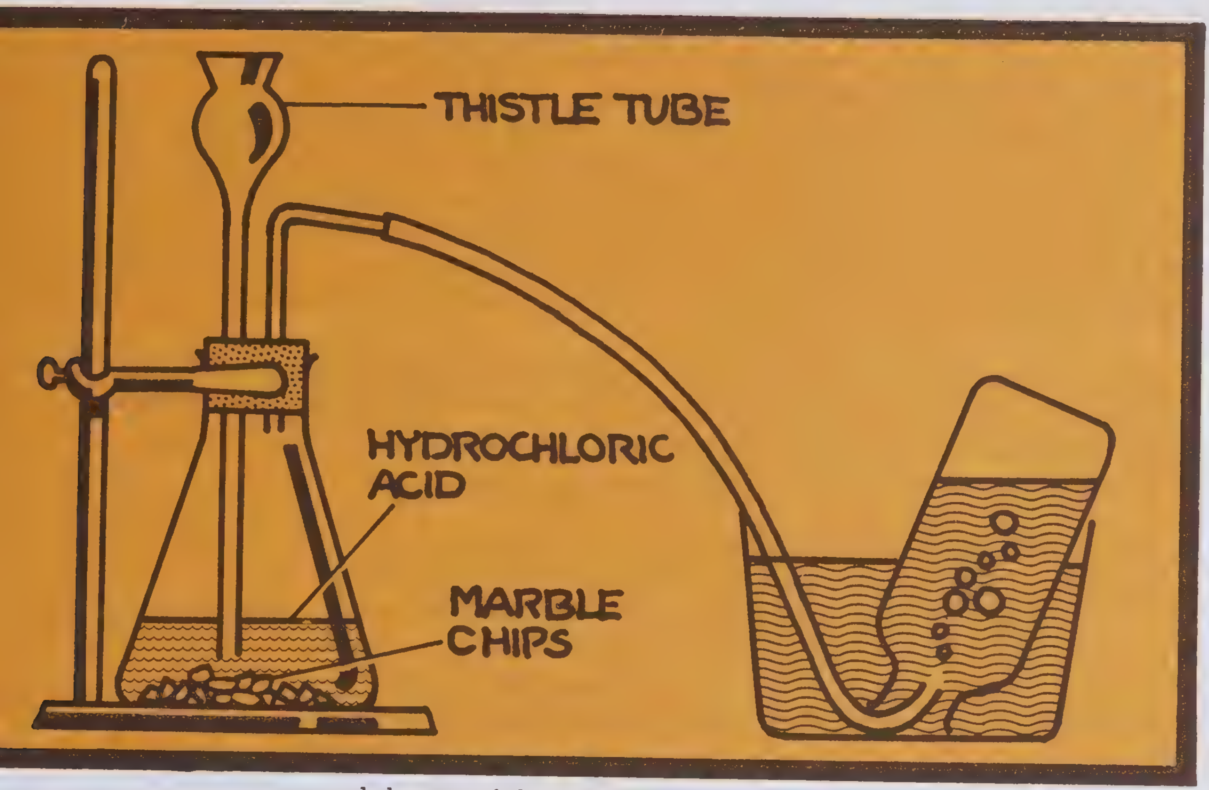
Preparation, Collection and Properties of Carbon Dioxide

retort stand and clamp
250 ml erlenmeyer flask
thistle tube
delivery tubing
stoppers
pneumatic trough
3 gas bottles
3 glass plates
splint
400 ml beaker
candle

125 ml erlenmeyer flask
retort ring
wire gauze
bunsen burner

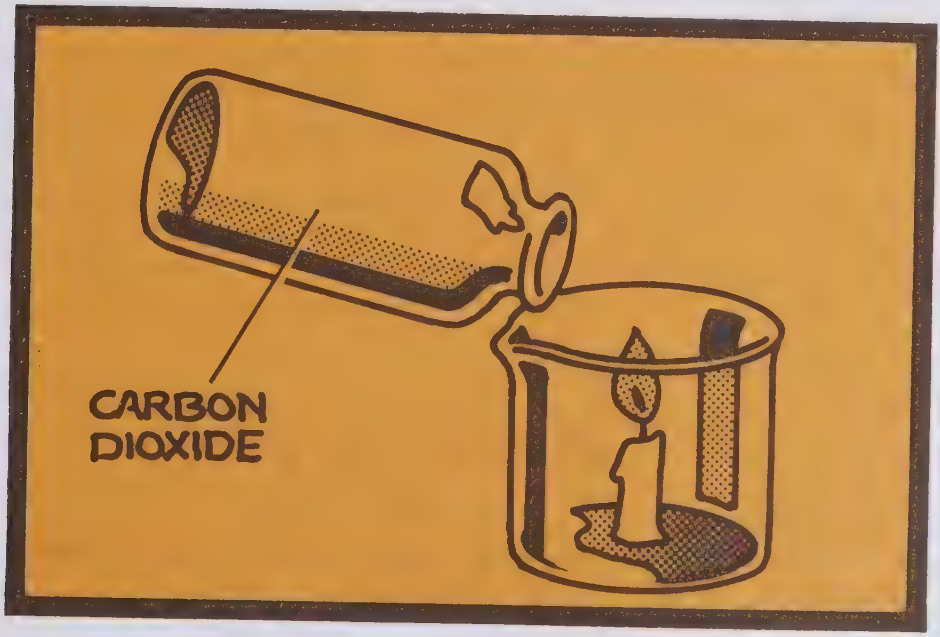
calcium carbonate (marble chips)
4 M HCl
neutral bromthymol blue
limewater
distilled water

1. Set up a gas generator similar to the one used for the preparation and collection of hydrogen. Use marble chips in the flask instead of zinc.
2. Invert three gas bottles full of water into a tray of water. Add 4 M HCl by the thistle tube until a vigorous reaction is taking place


THISTLE TUBE
HYDROCHLORIC ACID
MARBLE CHIPS

and the tip of the thistle tube is immersed in the liquid. Collect three bottles of the gas by the downward displacement of water.

3. Cover the three bottles of gas and turn them upright on your desk.
4. Observe the colour and odour of the gas in the first bottle, then thrust a burning splint into the bottle.
5. Add about 10 ml of distilled water to the second bottle and cover the mouth of the bottle with the palm of your hand. Shake the


CARBON DIOXIDE

bottle vigorously for several minutes. Note the sensation on your hand as you remove it from the mouth of the bottle. Add a few drops of neutral bromthymol blue to the solution in the bottle.

6. Light the stub of a candle and place it in the bottom of a 400 ml beaker. Pour the third bottle of gas into the beaker (see diagram).
7. If the reaction in the generator has stopped, add a little more acid solution. Place the end of the delivery tube into a 125 ml erlenmeyer flask that contains 20 ml of limewater. Pass the gas through the limewater until there appears to be no further change in it.
8. Place the flask and contents from Step 7 on a wire gauze supported on a retort ring and boil the liquid for a couple of minutes. Record all observations.

Questions

1. List the properties of carbon dioxide as determined in this experiment. State the colour, odour, solubility, ability to burn and ability to support combustion.
2. What sensation did you notice on the palm of your hand from Step 5? What caused this sensation? What type of solution was formed between the carbon dioxide and the water?
3. What three properties of carbon dioxide were illustrated by Step 6?
4. Calcium carbonate reacts with hydrochloric acid to produce calcium chloride, water and carbon dioxide. Write a balanced chemical equation for the reaction.
5. From the properties of carbon dioxide that you have discovered in this experiment what uses can you suggest for the gas?
6. Limewater is an aqueous solution of calcium hydroxide. Write the formula for calcium hydroxide.
7. Calcium hydroxide reacts with carbon dioxide to produce calcium carbonate. Write the equation for this reaction. What properties of calcium carbonate are illustrated in Step 7?
8. Calcium carbonate reacts with the carbon dioxide and water to produce calcium hydrogen carbonate. Write the equation for this reaction. Is calcium hydrogen carbonate soluble in water? How do you know? What happens to calcium hydrogen carbonate when it is heated?

Stoichiometry

The coefficients used to balance a chemical equation represent the number of molecules of each substance required for the Law of

Conservation of Mass to be obeyed. For example:

1. $NaOH + HCl \longrightarrow NaCl + H_2O$

is balanced and therefore obeys the Law of Conservation of Mass. However, the next three equations all obey the same law:

2. $2NaOH + 2HCl \longrightarrow 2NaCl + 2H_2O$
3. $10NaOH + 10NaCl \longrightarrow 10NaCl + 10H_2O$
4. $nNaOH + nHCl \longrightarrow nNaCl + nH_2O$

If "n" were Avagadro's number, 6×10^{23}, the equation would *still* be balanced. Since 6×10^{23} is a mole we could write the balanced equation.

5. $1 \text{ mol } NaOH + 1 \text{ mol } HCl \longrightarrow 1 \text{ mol } NaCl + 1 \text{ mol } H_2O$

The coefficients in this equation are the same as those in Equation 1. Since it is much easier to measure moles than individual molecules, chemists have reached an agreement among themselves.

> "The coefficients in a balanced equation represent *moles* of material. An unwritten coefficient is understood to be 1 mol."

With this understanding we are now back to:

6. $NaOH + HCl \longrightarrow H_2O + NaCl$

However now the equation is expressed in moles.

From the equation we can get *mole ratios*. For example:

$$\frac{\text{moles } H_2O}{\text{moles } NaCl} = \frac{1}{1}$$

$$\frac{\text{moles } NaOH}{\text{moles } H_2O} = \frac{1}{1}$$

$$\frac{\text{moles } NaOH}{\text{moles } HCl} = \frac{1}{1}$$

Consider the last example. The equation tells us that 1 mol of sodium hydroxide will be consumed in producing 1 mol of water.

Let's go one step further.

1 mol of sodium hydroxide has a mass of $23 + 16 + 1 = 40$ g

1 mol of water has a mass of $2 + 16 = 18$ g

Therefore 40 g of sodium hydroxide (1 mol) will produce just 18 g of water (1 mol).

Let's write the equation again showing the mass of each substance below it:

$NaOH$	$+ HCl$	$\longrightarrow NaCl$	$+ H_2O$
1 mol	1 mol	1 mol	1 mol
40 g	36.5 g	58.5 g	18 g

First of all, notice that the sum of the masses on each side of the reaction are equal — conservation of mass again.

$40 \text{ g} + 36.5 \text{ g} = 76.5 \text{ g} = 58.5 \text{ g} + 18 \text{ g}$

Notice too, that, for example:

36.5 g of hydrochloric acid are required to produce 58.5 g of sodium chloride.

If you like, expressed as a ratio:

$$\frac{\text{moles HCl}}{\text{moles NaCl}} = \frac{1}{1} \qquad \frac{\text{grams HCl}}{\text{grams NaCl}} = \frac{36.5}{58.5}$$

To a chemist, the important aspect of all this is that knowing the ratio of masses—which he obtained from the mole ratio—which he obtained from the balanced equation—he can calculate amounts for *any* quantity.

In this course we would rarely want to produce 58.5 g of sodium chloride. How much acid and how much base would be required to produce 5 g of NaCl?

Looking back we see that:

36.5 g of hydrochloric acid are required to produce 58.5 g of sodium chloride. So x g of hydrochloric acid will produce 5 g of sodium chloride.

The ratio is set:

$$\frac{36.5\text{ g}}{x\text{ g}} = \frac{58.5\text{ g}}{5\text{ g}}$$

Solving for x

$$x\text{ g} = \frac{36.5\text{ g} \times 5\text{ g}}{58.5\text{ g}}$$

$$= 3.12\text{ g}$$

Therefore we would require 3.12 g of acid to begin with.

Example 1

40 g sodium hydroxide produce 58.5 g sodium chloride

So x g sodium hydroxide produce 5 g sodium chloride

$$\frac{40\text{ g}}{x\text{ g}} = \frac{58.5\text{ g}}{5\text{ g}}$$

$$x\text{ g} = \frac{40\text{ g} \times 5\text{ g}}{58.5\text{ g}}$$

$$= 3.41\text{ g}$$

Therefore, looking back at our equation:

$$\begin{array}{lllll}
 & NaOH + & HCl \longrightarrow & NaCl + & H_2O \\
 & 1\text{ mol} & 1\text{ mol} & 1\text{ mol} & 1\text{ mol} \\
 & 40\text{ g} & 36.5\text{ g} & 58.5\text{ g} & 18\text{ g} \\
\text{And now} & 3.41\text{ g} & 3.12\text{ g} & 5\text{ g} &
\end{array}$$

We can get the amount of water by subtraction.

$$3.41\text{ g} + 3.12\text{ g} = 5\text{ g} + x\text{ g}$$
$$6.53\text{ g} = 5\text{ g} + x\text{ g}$$
$$1.53\text{ g} = x\text{ g}$$

Therefore, 1.5 g of water would be produced in these circumstances.

We took the long route with that example. The procedure can be condensed considerably if you follow a pattern. Let's look at another example.

Example 2

What mass of water would be produced by the neutralization of 50 g of sulphuric acid with potassium hydroxide?

1. Balance the equation.

$$H_2SO_4 + 2KOH \longrightarrow K_2SO_4 + 2H_2O$$

2. Write the masses represented by the equation *below* the symbols.

$$\begin{array}{lllll} H_2SO_4 + & 2KOH & \longrightarrow & K_2SO_4 + & 2H_2O \\ 98\text{ g} & 112\text{ g} & & 174\text{ g} & 36\text{ g} \end{array}$$

Check your masses by confirming conservation of mass.

$$98\text{ g} + 112\text{ g} = 210\text{ g} = 174\text{ g} + 36\text{ g}$$

You can often pick up a mistake in addition here!

Write the required information and the given data *above* the symbols.

$$\begin{array}{lllll} 50\text{ g} & & & & x\text{ g} \\ H_2SO_4 + & 2KOH & \longrightarrow & K_2SO_4 + & 2H_2O \\ 98\text{ g} & 112\text{ g} & & 174\text{ g} & 36\text{ g} \end{array}$$

Now set up the ratio:

50 g of sulphuric acid produce x g of water
98 g of sulphuric acid produce 36 g of water

$$\frac{50\text{ g}}{98\text{ g}} = \frac{x\text{ g}}{36\text{ g}}$$

$$x\text{ g} = \frac{50\text{ g} \times 36\text{ g}}{98\text{ g}}$$

$$= 18.4\text{ g} = 18\text{ g}$$

Therefore, 18 g of water should be produced.

That still looks like a long solution but only because we rewrote the equation several times. Look at this example and try to follow it.

Example 3

What mass of zinc chloride would be produced if 25 g of zinc were reacted with hydrochloric acid?

$$\begin{array}{lllll} 25\text{ g} & & & x\text{ g} & \\ Zn & + 2HCl & \longrightarrow & ZnCl_2 & + H_2 \\ 65.4\text{ g} & + 73\text{ g} & & 136.4\text{ g} & + 2\text{ g} \end{array}$$

Check for conservation of mass.

$$\frac{25\text{ g}}{65.4\text{ g}} = \frac{x\text{ g}}{136.4\text{ g}}$$

$$x\text{ g} = \frac{25\text{ g} \times 136.4\text{ g}}{65.4\text{ g}}$$

$$= 52.1\text{ g} = 52\text{ g}$$

Therefore 52 g of zinc chloride would be produced.

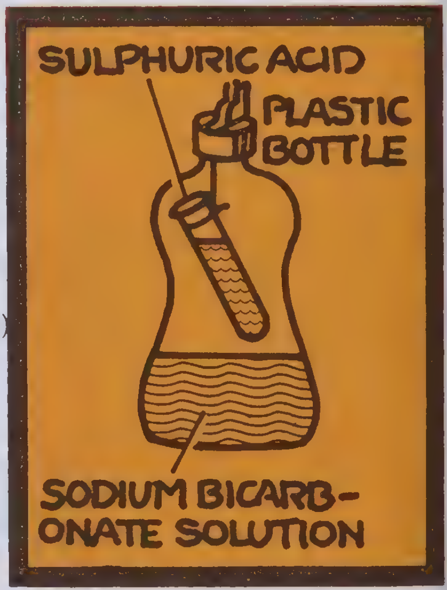
Problems

1. What mass of hydrogen peroxide will be required to produce 32 g of oxygen gas?
2. How much oxygen gas could be produced by the decomposition of 10 g of potassium chlorate? Express your answer first in grams, then in moles, then in litres at STP.
3. How much carbon would be required to produce 25 g of carbon dioxide gas?
4. What mass of potassium hydroxide will be needed exactly to neutralize 200 g of acetic acid?
5. You wish to produce 100 g of hydrogen gas in the laboratory. Which acid would be more economical?
   (a) Nitric acid at $1.00/kg.
   (b) Sulphuric acid at 50¢/kg.
   (c) Acetic acid at 25¢/kg.

4.21 A Soda-Acid Type Fire Extinguisher (Demonstration)

plastic squeeze bottle equipped with a valve that opens when pressure is applied
small test tube
wire

baking soda (sodium bicarbonate)
6 M H_2SO_4



1. Half fill the plastic squeeze bottle with water and add about a large scoopula full of baking soda (sodium bicrabonate). Shake the bottle well to mix the soda with the water.
2. Suspend a small test tube on a wire as illustrated in the diagram. The test tube should be filled with 6 M H_2SO_4. Stopper the bottle tightly with the valve-equipped lid. The fire extinguisher is ready for use.

Warning

Make sure the valve works before you try this experiment.

3. Invert it over a sink.

Questions

1. When the bottle was inverted the acid mixed with the water and reacted with the sodium bicarbonate to produce a gas. Where does this gas go? Why?
2. What happens to the pressure inside the bottle?
3. Why did the water spray out when the bottle was inverted?
4. When this type of extinguisher is used on a fire, what is responsible for actually putting out the fire?
5. Examine the diagram of a soda-acid fire extinguisher. Explain how it works.

4.22 Dry Ice

dry ice
balloon
gas bottle and stopper
splint
beaker
limewater
litmus paper
tongs
test tube
2 M HCl

Warning

Do not handle dry ice with your fingers. Use tongs or forceps.

1. Drop a small piece of dry ice (about 1 cm^3) into half a test tube of water. Add strips of red and blue litmus paper.
2. Drop a small piece of dry ice (1 cm^3) into half a test tube of limewater. Save the resulting solution for later.
3. Drop a small piece of dry ice into the balloon and quickly knot the neck of the balloon.
4. Drop a small piece of dry ice into a gas bottle. Wait a moment then test the gas in the bottle with a burning splint.
5. Drop several pieces of dry ice into a beaker that is half filled with warm water. When the reaction is finished and no dry ice remains, taste the water.
6. Pour about 5 ml of 2 M HCl into the solution that you saved from Step 2.

Questions

1. Why is this material called "dry ice"?
2. What is dry ice?
3. Why is dry ice such a popular coolant in stores and refrigerated trucks?

Summary

It is not our normal procedure to list information for you but carbon dioxide and the carbonate family are very useful substances that you should hear about. Both substances are of great importance to man.

You must know that carbon dioxide is produced in the body tissues as a product of the energy-producing reactions known as respiration. You must also know that carbon dioxide is taken in by plants as one of the raw materials in photosynthesis. The output of carbon dioxide by all living things during respiration is balanced by the intake of carbon dioxide by plants during photosynthesis. The amount of carbon dioxide in the atmosphere has remained fairly constant at approximately 0.03 percent over the time span that it has been measured.

The amount of carbon dioxide in the air is critical to atmospheric conditions. A small increase in the amount of carbon dioxide could result in a general increase in the atmosphere's temperature. This could easily be followed by a melting of polar icecaps and widespread coastal flooding as the level of the oceans began to rise. This is one reason why we must monitor atmospheric carbon dioxide levels.

Carbon dioxide is slightly soluble in water and so is found in all natural waters. In this way aquatic plants are able to obtain the carbon dioxide they need to carry on photosynthesis.

This same solubility of carbon dioxide is used to carbonate beverages such as soft drinks, beers and sparkling wines. Champagne is a self-carbonated white wine produced when the bottle is sealed before the fermentation is complete. Gas produced by the fermentation is dissolved in the wine itself.

It was Joseph Priestley who is credited with discovering the "pleasant and rather tart taste" of carbon dioxide dissolved in water. Priestley called the solution "seltzer" water. Think of the names of commercial medicines that use that word! Today millions of kilograms of carbon dioxide are used every year to produce the soda water the world has come to love so much.

Carbon dioxide does not burn and does not support ordinary combustion. It is also denser than air. These properties make it ideal as a gas for extinguishing fires. Many homes, schools and industries have carbon dioxide fire extinguishers. Your chemistry laboratory should have a fire extinguisher in it. Check to see if it is a carbon dioxide cylinder.

The reaction of carbon dioxide and water to produce carbonic acid provides the parent acid for the carbonate family. Large quantities of the earth are made up of carbonates of one sort or another.

Perhaps the most plentiful of these is calcium carbonate or limestone, laid down on the floors of ancient oceans millions of years ago. These sedimentary rocks were produced as the calcium-carbonate-rich shells of sea animals gradually accumulated into thick layers on the ocean floors and then were compressed by the great masses of silt and water above them. A hammer and chisel can almost always pry loose fossils of sea animals from limestone.

Calcium carbonate or limestone can be heated in a furnace or kiln to produce lime:

$$CaCO_3 \longrightarrow CaO + CO_2$$

Lime has been used in the construction industry for thousands of years. Calcium oxide or lime is mixed with sand and water to form mortar. Lime and water mixed to form a thick paste makes plaster. Lime and water mixed to form a soupy liquid makes whitewash. Lime and water mixed to form a saturated solution forms limewater.

In all these applications where lime is mixed with water the typical metallic oxide and water reaction takes place to form a base:

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Carbon dioxide in the air then reacts with this base to form the carbonate again:

$$Ca\,(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

Pure calcium carbonate is a clean white solid. You have seen it many times in the form of blackboard chalk. The famous White Cliffs of Dover are chalk cliffs.

If chalk or limestone is compressed and heated deep in the earth's crust a crystalline form of calcium carbonate called marble is formed. Mineral impurities often give marble a colour or streaked appearance that gives it great beauty when polished. Many great works of art such as Michaelangelo's Moses and Pieta have been carved in marble. Many public buildings have used marble extensively on floors and walls.

Sodium carbonate (Na_2CO_3) or washing soda and sodium bicarbonate ($NaHCO_3$) or baking soda are used in many homes.

Copper carbonate ($CuCO_3$) or malachite is a common source of metallic copper.

Magnesium carbonate, often combined with calcium carbonate, is a common cause of hardness in water.

We could go on and list many more applications that have been found for these substances. Our intention here was simply to make you aware that many substances used in the chemistry laboratory have sound practical and commercial uses. It is the role of the chemist to discover new substances and to determine the properties of these substances. Once this is done engineers work out the technology for preparing these substances in quantity and quality. At this point the marketplace takes over as these substances are put to commercial use.

Many familiar substances such as lime have been around for a long time but most of the substances that you are in contact with every day have been produced in quantity only in the last 100 years or so.

Thinking of the future of chemistry is a job for someone with a real imagination!

Questions

1. Identify three different methods of preparing oxygen in the laboratory. Write an equation to represent each method.
2. What is a catalyst?
3. It has been suggested that copper (II) sulphate will catalyse the reaction between zinc and hydrochloric acid. Describe how you would verify this theory.
4. You are given three unlabelled bottles of gases which are all clear and colourless. You are told one is oxygen, one is hydrogen and one is carbon dioxide. Describe how you would determine which bottle contains which gas.
5. What is an indicator?
6. Oxygen can be collected by the downward displacement of water. What properties must a gas have before this method can be used to collect it?
7. Compare those factors that tend to increase the level of atmospheric carbon dioxide with those factors that tend to decrease it.
8. Suggest a reason why hydrogen is not found in the atmosphere as hydrogen gas.
9. Describe how you would attempt to determine if there is any hydrogen present in a sample of wood or newspaper.
10. On returning from the Christmas holiday you find a bottle of a clear colourless liquid left in your desk unlabelled. You do not remember what it is but think it may be plain water. Describe how you would determine if the bottle contains:
    (a) plain water;
    (b) a water solution of something;
    (c) some other substance.
11. Pure hydrochloric acid is 37 percent by mass hydrogen chloride. What volume of chlorine gas must be used to prepare 1 ℓ of acid?
12. In the future, as petroleum supplies dwindle we may return to the water gas reaction. In this reaction hot charcoal reacts with water to produce a gaseous fuel called water gas:
    $$C + H_2O \longrightarrow CO + H_2$$
    Both products are combustible:
    $$2\,CO + O_2 \longrightarrow 2CO_2$$
    $$2\,H_2 + O_2 \longrightarrow 2H_2O$$
    Calculate the volume of oxygen at 25°C and 750 mm pressure required to burn the water gas produced when 100 kg of carbon

are converted by excess water into the mixture of carbon monoxide and hydrogen.

13. Acetylene gas (used in welding) can be prepared as follows:

    $$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$$

    What mass of calcium carbide would be needed to produce enough gas to fill a large metal cylinder of 100 ℓ capacity if it is at 2500 kPa pressure and 10°C?
14. Which substance could produce the greater volume of oxygen?
    (a) 100 g of mercury (II) oxide;
    (b) 25 g of potassium chlorate;
    (c) 10 g of hydrogen peroxide.
15. The coal used in electricity generating stations is about 85 per cent carbon. The flue gases are at about 500°C and 800 mm pressure. What volume of gas passes out of the stack of a furnace that consumes 800 kg of coal every hour? Assume that air is 20 percent oxygen and that only half the input oxygen is converted to carbon dioxide. Assume no other gases are produced.
16. What volume would 5000 g of sulphur hexafluoride (SF_6) occupy at 200°C and 800 mm pressure?
17. Calculate the molarity of a solution of sodium chloride if 50 g of solute are dissolved and diluted to make 1 ℓ of solution.
18. What volume of 0.3 M $Ca(OH)_2$ would be required to neutralize 500 ml of 1.5 M HNO_3 solution?
19. 50 ml of 0.1 M NH_4OH are added to 100 ml of 0.2 M HCl. The resulting solution is evaporated to dryness. Calculate the mass of the dry residue.
20. How many grams of a 20 percent solution of nitric acid would be required to neutralize 100 ml of 0.1 M $Al(OH)_3$?

Problems

1. How many moles of copper are necessary to produce 5 mol of copper (II) sulphide?
2. What would be the minimum number of moles of oxygen consumed by burning 1 mol of phosphorus?
3. How many moles of phosphoric acid would be neutralized by 10 mol of magnesium hydroxide?
4. How many moles of hydrogen peroxide would be needed to produce the oxygen necessary to form 3 mol of aluminum oxide?
5. How many moles of water would be formed if 2 mol of hydrogen gas were reacted with 2 mol of oxygen gas?
6. How many moles of magnesium oxide are present in a 100 g sample of that substance?
7. How many grams of sulphur would be required to produce 5 mol of sulphur (IV) oxide?
8. What mass of potassium chlorate would be needed to produce 0.5 mol of oxygen gas?

9. Two moles of hydrochloric acid will produce what mass of hydrogen gas when it is reacted with:
   (a) zinc;
   (b) aluminum;
   (c) iron?
10. How many grams of sodium hydroxide will be needed to neutralize 100 ml of 0.1 M HCl that has been spilled on a table top?
11. What mass of aluminum is needed to react completely with 100 g of liquid bromine?
12. How much mercury would be produced if 100 g of mercury (II) oxide were decomposed?
13. How much tin (II) fluoride could be produced from 100 g of pure tin?
14. What mass of hydrogen chloride would need to be electrolyzed in order to produce 0.5 kg of hydrogen?
15. Silicon carbide is the second hardest substance known to man. What mass of silicon would be needed to produce 100 kg of silicon carbide?
16. What mass of zinc would be needed to produce 500 ml (STP) of hydrogen from hydrochloric acid?
17. What mass of water would be needed to produce 1000 ℓ (STP) of oxygen gas by electrolysis?
18. What mass of sulphur could be converted into sulphur (VI) oxide by 1 ℓ of oxygen gas at STP?
19. What volume of carbon dioxide would be produced if a 7 kg bottle of nitric acid were poured over a large mass of limestone?
20. What volume of oxygen would be needed completely to oxidize 5 kg of barbeque charcoal?
21. How many litres of oxygen will be needed to react completely with 5 ℓ of hydrogen gas — both at STP?
22. Hydrogen gas combines with chlorine gas to form hydrogen chloride gas. What volume of hydrogen chloride gas could be produced from:
    (a) 100 ℓ of hydrogen at STP;
    (b) 5 ℓ of chlorine at STP?
23. Ammonia gas is prepared by reacting hydrogen and nitrogen together:

    $$N_2 + 3H_2 \longrightarrow 2NH_3$$

    Give the volume of each reactant needed to produce 1000 ℓ of ammonia at STP.
24. Propane, C_3H_8, reacts with oxygen to form *only* water and carbon dioxide. What volume of oxygen is required to burn completely 100 ℓ of propane gas at STP?
25. What volume of carbon dioxide would be formed if 25 ℓ of oxygen were used to burn a sample of pure carbon?
26. A 100 g sample of pure iron is dropped into sulphuric acid. What volume of hydrogen measured at 20°C and 700 mm would be produced?
27. What volume of carbon dioxide gas measured at − 20°C and

800 mm pressure could be produced by reacting 100 g of sodium carbontate with an acid?

28. The ammonia produced by the combination of nitrogen gas and hydrogen in the Bosch process is at about 1000°C and 5000 kPa of pressure. What volume of ammonia would be produced from 1000 ℓ of hydrogen at 25°C and 760 mm pressure?
29. 10 ℓ of hydrogen at 25°C and 760 mm are mixed with 6 ℓ of oxygen at 20°C and 740 mm pressure. The resulting mixture is exploded. What mass of water would be produced?
30. What volume of oxygen measured at 25°C and 800 mm pressure would be available from 5 mol of hydrogen peroxide at 20°C and 760 mm pressure?

Some Words You Should Know

catalyst
catalysis
peroxide
prefix
inhibitor
enzyme
acid
base
operational definition
conceptual definition
hydroxide
organic
oxy-acid
binary
ternary
strong acid
weak acid
hydronium
neutralization
neutral
titration
molarity
burette
meniscus
displacement
radical
ion
equation
balance
oxidation
reduction
limewater
stoichiometry
carbonate
empirical
formula

5

Active Metals/The Halogens/Inert Gases/The Periodic Table/Valence Numbers/Covalent/Atomic Diameter

The Active Metals

So far in this course we have looked at elements and compounds and the relationships between the two. The time has come to try and see how the various pieces of the picture fit together. We need a broader perspective, an overview and a focal point.

Many scientists feel that the greatest achievement in chemistry belongs to Dmitri Mendeleev (1834-1907) the great Russian chemist who provided a focal point for the scientists of the late nineteenth century.

In this unit, we hope to share Mendeleev's ideas with you. First, as always, we will do some experiments. We cannot go anywhere in science without the facts!

5.1

Reaction of Metals With Water

calcium
magnesium
lithium
sodium
potassium
phenolpthalein solution
glass plate
tongs
2 250 ml beakers
2 test tubes
splints
tweezers
knife
drinking straw

3 1 **Li** 6.9	
11 1 **Na** 23.0	12 2 **Mg** 24.3
19 1 **K** 39.1	20 2 **Ca** 40.1

Warning

Handle all metals with tongs.

1. Carefully examine a small piece of each of the metals listed above. As part of your examination scrape the surface of the metal with the knife. Note the nature and colour of the freshly exposed surface and the speed with which the surface becomes discoloured.

Warning

Do not touch lithium, sodium or potassium with your fingers.

Be sure that after your examination lithium, sodium and potassium are again immersed in kerosene or mineral oil.

Part A Calcium and Magnesium

1. Fill two test tubes with water. Invert them both in a 250 ml beaker half filled with water.
2. Drop in a lump of calcium that has had a small area scraped clean of the surface oxide. Collect the gas produced in the two test tubes.
3. Remove the two test tubes filled with gas from the water. Hold one test tube upright and the second inverted for one full minute. Bring a burning splint to the mouth of each test tube.
4. Divide the liquid in the beaker into two parts. Filter one part. Add five drops of colourless phenolphthalein to the second part.
5. Using a drinking straw, exhale through the filtrate of the first part for several minutes.
6. Heat half a test tube of distilled water to boiling. Remove from heat and add a 3 cm strip of freshly cleaned magnesium ribbon. Observe closely. (The magnesium should be cleaned with fine steel wool.)
7. Add three drops of colourless phenolphthalein to the test tube.

Questions

1. From your examination of the metals in Part A, what would you say these metals have in common? How do they differ?
2. Why do you think some of these metals are stored in kerosene?
3. What gas is produced by the reaction of calcium with water?
4. List four properties of hydrogen illustrated by this experiment.
5. What substance was identified as the distillate in Steps 4 and 5?
6. Write a balanced chemical equation for the reaction between calcium and water.
7. On the basis of the very limited evidence collected write an equation for the reaction between magnesium and water.
8. Which of these metals is more active in its reaction with water?

Part B Lithium, Sodium and Potassium

Warning

Exercise extreme caution with these elements.

☞

1. Obtain a small piece of lithium (not larger than 0.5 cm^3). Dry off the excess oil with a paper towel. Drop the metal into a 250 ml beaker that is one-quarter filled with water. Cover the beaker *immediately* with the glass plate. You should look for the following:
    (a) the shape of the lithium;
    (b) the movement of the lithium;
    (c) the relationship between the lithium and the surface of the water;
    (d) any materials produced by the reaction.
2. After the reaction has ceased, add five drops of colourless phenolphthalein to the beaker.
3. Rinse out the beaker and repeat Steps 1 and 2 with sodium and finally with potassium.

Questions

1. Are these exothermic or endothermic reactions?
2. What substance is produced in each case?
3. What other substance do you suspect is produced by the reaction?
4. Write equations for each reaction.
5. List these three metals in order of the activity of their reaction with water.
6. Add calcium and magnesium in their proper places to the list from Question 5.
7. Write a general equation to represent the action of all five metals with water.
8. How do the properties of these metals compare?
9. Find these metals in the periodic table. How do these metals come to be placed where they are?
10. Make some predictions about the behaviour of cesium and barium.

The Halogens

5.2a

To Prepare and Study Chlorine

sodium chloride
manganese dioxide
concentrated sulphuric acid
litmus paper
test tubes
clamp
bunsen burner
watch glass
splint

17 1 5 7
Cl
35.5

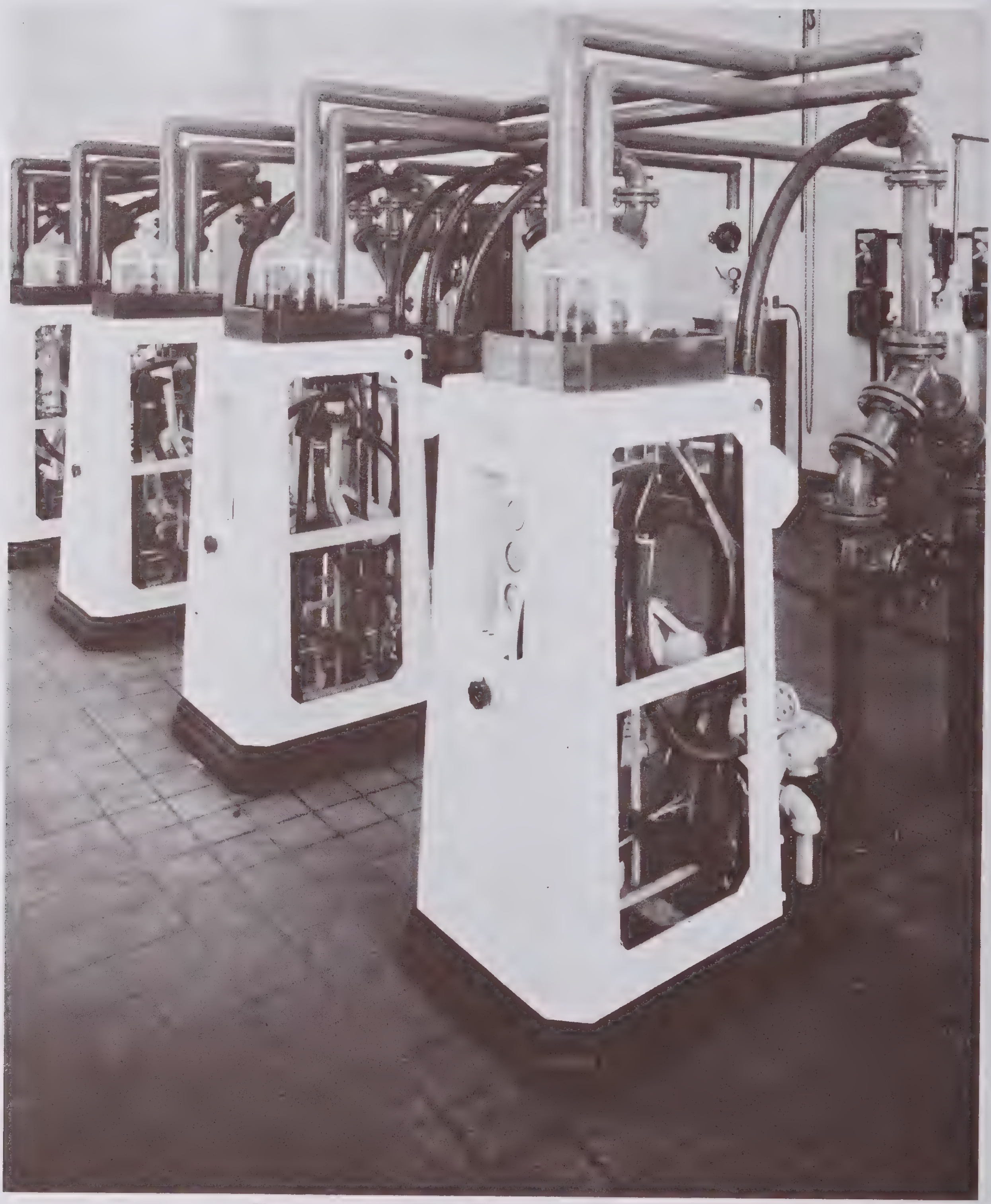
Warning

The chlorine gas produced in this experiment is poisonous. Do not inhale the gas.

1. Place equal small quantities (enough to cover the end of your scoopula) of sodium chloride and manganese dioxide on a watch glass and mix the two thoroughly. Using a clean scoopula transfer the mixture to the bottom of a clean dry test tube. Be very careful not to get the mixture onto the sides of the tube. Holding the test tube with a clamp add seven drops of concentrated sulphuric acid. With the tube near the exhaust vent, warm it *gently* with your bunsen burner for about a minute or until a slight effervescence appears in the test tube.
2. Attempt to determine the colour of the gas. It may help to hold the test tube up against a white background.
3. Note the odour of the gas. (See Unit 8 for special instructions.)
4. Lower strips of moist red and blue litmus into the test tube. Record all observations.
5. Test the combustibility of the gas with a burning splint.

Questions

1. Prepare a summary of the properties of chlorine shown in this experiment.
2. If the products of the reaction are chlorine, sodium sulphate, manganese sulphate and water, write a balanced chemical equation for the reaction.



Chlorinator unit. Units such as this add small amounts of chlorine to a city's water supply in order to purify the water. *Courtesy of Metropolitan Toronto, Department of Works.*

5.2b

Demonstration of Additional Properties of Chlorine

bottled chlorine
copper foil
powdered antimony
strips of cloth
tongs
bunsen burner
250 ml beaker

1. Obtain a strip of copper foil 1 cm by 5 cm. Hold it with a pair of tongs and heat it in a bunsen flame. Immediately lower the copper into a gas bottle of chlorine. When the reaction ceases let the residue drop to the bottom of the bottle and add 1 cm of distilled water. Shake well. Observe.
2. Sprinkle powdered antimony into a second bottle of chlorine. Observe.
3. Invert a bottle of chlorine into a large beaker of water to test the solubility of the gas.
4. Add strips of moist coloured cloth to another bottle of chlorine. Observe the effect after several minutes.

Questions

1. List the additional properties of chlorine shown in this part of the experiment.
2. Write balanced chemical equations where reactions occurred. This may require the use of a reference book.
3. What do the elements copper and antimony have in common? What general group of elements are they representative of?

5.3a

To Prepare and Study Bromine

sodium bromide
manganese dioxide
concentrated sulphuric acid
test tubes
clamp
bunsen burner
litmus paper
carbon tetrachloride
watch glass
2 test tubes
splint

35 1 5
Br
79.9

1. Place equal small quantities of sodium bromide and manganese dioxide on a watch glass and mix thoroughly. Using a clean spatula transfer the mixture to a dry test tube. Be careful not to get the mixture onto the sides of the test tube. Add 10 drops of concentrated sulphuric acid and warm the mixture gently for about a minute.
2. Note and record the colour and odour of the vapour that comes off from the mixture.

Warning

Be very careful not to breathe bromine. It is poisonous.

3. Test the vapour with strips of moist red and blue litmus paper.
4. Decant some of the vapour into a test tube containing 2 or 3 cm of water. Stopper the test tube containing the water and shake. Note the colour of the water. Add 1 ml of carbon tetrachloride to this test tube and shake again. Note the colour in the carbon tetrachloride layer.
5. Test the vapour with a burning splint.

Questions

1. Write a balanced chemical equation for the reaction that occurred in the original test tube.
2. Summarize the properties of bromine shown in this experiment. Compare them to chlorine pointing out similarities or differences.

5.3b Demonstration

liquid bromine
copper foil
powdered antimony
strips of cloth
3 gas bottles
3 glass plates
bunsen burner

Warning

Liquid bromine is very corrosive.

1. Place three drops of liquid bromine in the bottom of three different gas bottles. Cover each with a glass plate and shake each bottle until the liquid vaporizes.
2. Heat a strip of copper foil (1 cm by 5 cm) with a bunsen burner then quickly lower it into one of the bottles of bromine. After a couple of minutes remove the foil and examine the surface for evidence of a chemical change.
3. Heat some powdered antimony on the end of a scoopula, then sprinkle it into a bottle of bromine. Record all observations.
4. Lower strips of moist coloured cloth into the third bottle of bromine. After a couple of minutes remove the cloth and wash it under a stream of water. Record all changes.

Questions

1. Was there evidence of a chemical reaction between the copper and the bromine? What indications were there?
2. Did the antimony powder react with the bromine? How do you know?
3. Was the bromine capable of bleaching the cloth as the chlorine did?
4. Write balanced chemical equations for the reactions that occurred.

5.4 To Prepare and Study Iodine

sodium iodide
manganese dioxide
concentrated sulphuric acid
ice
water
powdered antimony
mortar and pestle
bunsen burner
retort stand
test tube
litmus paper
strips of cloth
250 ml beaker
evaporating dish
carbon tetrachloride

53 $^{1}_{5}$
I
126.9

Warning
Do not inhale the iodine vapour produced.

1. Crush a good sized scoopula of sodium iodide with a mortar and pestle. Transfer this and an equal amount of manganese dioxide to a clean dry 250 ml beaker and mix thoroughly. Add sufficient concentrated sulphuric acid just to moisten the surface (about 20 to 25 drops). Quickly cover the beaker with an evaporating dish containing water and crushed ice (see diagram). Place the entire assembly on your retort stand and heat gently. The heating must be done *extremely* carefully and the evaporating dish should be kept cold for a successful experiment. Continue the gentle heating for about 5 min after coloured fumes begin to form. Stop heating and allow the beaker to cool for a couple of minutes. Remove the evaporating dish and examine the underside for iodine.
2. Place a few of the crystals from the evaporating dish into the bottom of a clean test tube. Hold the tube at a 45° angle and heat gently until the tube fills with vapour. Lower strips of moist red and blue litmus into the vapour. Lower a moist strip of coloured cloth into the vapour for a moment. Remove the cloth and wash well. Record all observations. Sprinkle a little heated antimony powder into the iodine vapour.
3. Test the solubility of iodine crystals in (a) water (b) carbon tetrachloride. Use the iodine crystals that you prepared in Step 1.

Note: Iodine vapour will crystallize in your lungs as it did on the watch glass. Keep your nose away from the reaction test tube.

Questions

1. Write a balanced chemical equation for the preparation of iodine.
2. Does iodine:
   (a) react with a metal;
   (b) bleach cloth;
   (c) react with litmus?
3. What change of state does the iodine undergo when heated *gently*? Why must the heating be done gently?

Now consider chlorine, bromine and iodine together.

1. Find these elements in the periodic table. Why do you think they are where they are?
2. Can you make any predictions about fluorine?

5.5 Relative Activity of the Halogens

sodium chloride solution
sodium bromide solution
sodium iodide solution
chlorine water (fresh)
bromine water
3 test tubes
carbon tetrachloride

1. Place 2 to 3 ml of sodium iodide solution in each of two test tubes. Add 1 ml of carbon tetrachloride, cork the test tubes and shake well. Record your observations. To one of the test tubes add 1 ml of chlorine water. Cork the tube and shake well. Record your observations. To the second tube add 1 ml of bromine water. Cork the tube and shake well. Record your observations.
2. Place 2 to 3 ml of sodium bromide solution in a test tube. Add 1 ml of carbon tetrachloride to it, cork the test tube and shake well. Record your observations. Add 1 ml of chlorine water and shake well. Record your observations.

Questions

1. What colour is:
   (a) sodium iodide solution;
   (b) sodium bromide solution;
   (c) sodium chloride solution;
   (d) chlorine water;
   (e) bromine water?
2. What colour is:
   (a) bromine in carbon tetrachloride;
   (b) iodine in carbon tetrachloride?
3. Complete Table 5.1 by putting an x where a reaction occurred in this experiment.

Table 5.1

	Chloride	*Bromide*	*Iodide*
Chlorine			
Bromine			
Iodine			

4. Write the correct formula for:
   (a) sodium chloride
   (b) sodium bromide
   (c) sodium iodide
   (d) chlorine gas
   (e) bromine
   (f) iodine
5. One of the halogens used is capable of displacing each of the other two from their salts. Which halogen is it?
6. One of the halogens used is capable of displacing one of the other halogens from its salt. Which one is it?
7. Write balanced chemical equations for the reactions that occurred in this experiment.
8. The reactions that took place in this experiment were oxidation-reduction reactions. Pick out:
   (a) the substance oxidized in each reaction;
   (b) the oxidizing agent in each reaction.
9. What role was played by the carbon tetrachloride in this experiment?
10. Compare the three halogens studied by completing Table 5.2.

Table 5.2

	Chlorine	*Bromine*	*Iodine*
Atomic number			
Atomic mass			
Valence			
Electrons in valence shell			
Physical state (room conditions)			
Colour			
Odour			
Reaction with metals (rate each)			
Bleaching action			
Strength as an oxidizing agent			

11. Refer to a periodic chart and find chlorine, bromine and iodine on the table. In the same group you will find two other halogens. Name them.
12. On the basis of the information about halogens gathered so far predict the following for the two halogens that you have not studied:
    (a) physical state;
    (b) colour;
    (c) odour;
    (d) reaction with metals;
    (e) bleaching action;
    (f) strength as oxidizing agent.

5.6

Reaction of Sulphuric Acid with Sodium Chloride, Sodium Bromide and Sodium Iodide

Part A Sodium Chloride

sodium chloride
concentrated sulphuric acid
hydrochloric acid
ammonium hydroxide
litmus paper
beaker
2 test tubes
filter paper

1. Place enough sodium chloride to cover the end of your spatula in the bottom of a clean dry test tube. Add three or four drops of concentrated sulphuric acid until an effervescence occurs in the tube.
2. Hold moist strips of red and blue litmus paper in the mouth of the test tube. Observe.
3. Dip a small strip of filter paper into a bottle of ammonium hydroxide and just moisten the tip. Now bring the filter paper to the mouth of the test tube.
4. Note the colour of the gas and also the odour. (See Unit 8 for procedure.)
5. Add another drop or two of acid to increase the effervescence in the tube. After a moment quickly invert the entire test tube into a beaker of water. Note the solubility of the gas in water.
6. Repeat Steps 2 to 4 above only use a couple of drops of hydrochloric acid in place of the salt and sulphuric acid.

Questions

1. Summarize the properties of the gas produced under colour, odour, action on litmus, solubility and action with ammonia.
2. How did the gas from hydrochloric acid compare in properties to the gas produced in this experiment?
3. Write an equation for (a) the reaction between sodium chloride and sulphuric acid and (b) the reaction between the gas produced and ammonia.
4. What gas was produced when sulphuric acid reacted with sodium chloride?

Part B Sodium Bromide

sodium bromide
concentrated sulphuric acid
ammonium hydroxide
litmus paper
beaker
2 test tubes
filter paper
carbon tetrachloride

1. Place enough sodium bromide to cover the end of your spatula in the bottom of a clean dry test tube. Add three or four drops of sulphuric acid until an effervescence occurs in the test tube.
2. Hold strips of moist red and blue litmus paper in the mouth of the tube for about 30 s. Remove them and rinse them under the tap. Record your observations.
3. Dip a small strip of filter paper into a bottle of ammonium hydroxide and just moisten the tip. Now bring the filter paper to the mouth of the test tube.
4. Note the colour and the odour of the gas(es) (cautiously).
5. Decant some of the gas from the test tube into a second test tube containing 1 ml of carbon tetrachloride. Stopper the second test tube and shake well. Note and record the colour in the carbon tetrachloride layer.
6. Add a bit more sodium bromide and acid to increase the rate of bubbling in the test tube. After a minute quickly invert the test tube in a beaker of water to test the solubility of the product(s) of the reaction.

Questions

1. Assuming that a reaction similar to that in Part A occurred, what was one of the products of this reaction?
2. Summarize the properties of this product under colour, odour, action on litmus, solubility and action with ammonia.
3. Was there any indication that another gas was produced other than the one mentioned in Question 1?
4. Describe in as much detail as possible how the reaction in this part of the experiment differed from the reaction that occurred in Part A.
5. Write a balanced chemical equation for the reaction between sulphuric acid and sodium bromide.

Part C Sodium Iodide

concentrated sulphuric acid
sodium iodide
ammonium hydroxide
litmus paper
beaker
test tube
filter paper

1. Place enough sodium iodide to cover the end of your spatula in the bottom of a clean dry test tube. Add three or four drops of concentrated sulphuric acid until an effervescence occurs in the bottom of the test tube.
2. Hold strips of moist red and blue litmus paper in the mouth of the test tube for a minute. Remove the litmus and rinse it under the tap.
3. Hold a strip of filter paper that has been dipped into some ammonium hydroxide near the mouth of the test tube.
4. Cautiously note the odour(s) and the colour.
5. Increase the rate of bubbling in the test tube by adding a bit more sodium iodide and a couple more drops of acid. After a minute quickly invert the test tube into a beaker of water.

Questions

1. Assuming a reaction similar to Part A of this experiment, what gas should have been produced?
2. List as much evidence as you can to indicate that other reactions did occur producing substances other than the product desired.
3. Write an equation for the primary reaction between sodium iodide and sulphuric acid.
4. A secondary reaction occurred between the gaseous product of the primary reaction and sulphuric acid. Use a resource book from the library to find out what this secondary reaction was and write a balanced chemical equation for the reaction.
5. The primary product of each part of this experiment is the hydrogen *halide*. Name the three halides, describe what is meant by a halide and state why the method is only really suitable for the production of one of the halides.
6. Suggest a method for producing hydrogen fluoride.

These last experiments contain a great wealth of information about the halogens. If you try to remember the individual observations and facts you will probably be overwhelmed by it all. Instead you should be looking for generalizations and regularities. Similarities and differences will be a great help in finding these regularities.

This last experiment in the series will help you to tie many of the pieces together. Remember—look for the regularities! The hardest part of this experiment is trying to distinguish between the halogen and the halide, for example, chlorine and chloride. Take it slow and easy. Make sure you are aware of what you are dealing with and whether it is the atom or the ion!

5.7

Test for the Chloride, Bromide and Iodide Ions

silver nitrate
carbon tetrachloride
chlorine water
samples of a soluble chloride, bromide and iodide
aqueous ammonia
nitric acid
test tubes

Part A Chloride

1. The test for a soluble chloride depends on the insolubility of silver chloride. To a sample of a soluble chloride solution add a few drops of silver nitrate solution. Note and describe the product of this reaction.
2. Test the solubility of the precipitate by adding several millilitres of aqueous ammonia to the test tube.
3. Add a few drops of nitric acid to the test tube.

Questions

1. The precipitate is silver chloride. Write the balanced *ionic* equation for the reaction.
2. The precipitate reacts with ammonia to produce a complex silver ammonia ion, $Ag(NH_3)_2{}^+$. Write the equation for this reaction.

Part B Bromide

1. Bromine is very soluble in carbon tetrachloride. When it dissolves it gives a very distinctive colour to the carbon tetrachloride layer. Recall what this colour is.
2. To a sample of a soluble bromide add 2 to 3 ml of carbon tetrachloride. Shake the tube vigorously then allow it to settle for a moment. Now add a couple of millilitres of fresh chlorine water and again shake well. Note and record the colour in the carbon tetrachloride layer.

Questions

1. What role is played by the carbon tetrachloride layer? Does it enter into the reaction that occurs?
2. What role is played by the chlorine water? Use the proper terminology.
3. Write a balanced equation for the reaction.

Part C Iodide

1. Like bromine, iodine also gives a distinctive colour to carbon tetrachloride when dissolved in it. Recall what this colour is.
2. To a sample of a soluble iodide add 2 to 3 ml of carbon tetrachloride. Shake the tube vigorously then allow it to settle for a moment. Now add 2 to 3 ml of fresh chlorine water and again shake well. Note and record the colour in the carbon tetrachloride layer.

Questions

1. What role is played by the carbon tetrachloride?
2. What role is played by the chlorine water?
3. Write a balanced chemical equation for the reaction between chlorine and the iodide.
4. Which of these three halogens has the greatest affinity for electrons?
5. Elements that accept electrons to become negative ions are said to be *electronegative*. Which halogen is the most electronegative?
6. Name some other elements that are electronegative.
7. What would *electropositive* mean? Name some electropositive elements.

Building the Periodic Table

Let us consider for a moment some of the data we have collected in the last few experiments.

Sodium, potassium, lithium, calcium and magnesium all react with water to produce a gas. Tests with calcium show the gas is hydrogen. The second product in each case is a base. This leads us to the general equation:

$$\text{metal} + \text{water} \longrightarrow \text{hydrogen} + \text{base}$$

We must recognize that not all metals do this. Copper, lead and stainless steel are only three examples of metals that are often in contact with water, but do not react with it in this way. Let's call those metals that do obey our generalization "active" metals. Our equation becomes

$$\text{active metal} + \text{water} \longrightarrow \text{hydrogen} + \text{base}$$

The five active metals we have seen are sodium, potassium, lithium, calcium and magnesium. They differ in several ways. First, they all react with water at different speeds or rates.

Second, lithium, sodium and potassium are stored in kerosene or liquid paraffin because they react so readily with air.

Thirdly, lithium, sodium and potassium have valences of +1 while calcium and magnesium have valences of +2.

In fact, the similarities among lithium, sodium and potassium are many. This leads us to group them together into a family of elements. Within the family properties change — but in a regular way.

Table 5.3

Atomic Diameter	*(10* $^{-12}$ m)	MP(°C)
Lithium	313	186
Sodium	383	98
Potassium	476	63

We might also ask about rubidium, cesium and francium.

Table 5.4

Atomic Diameter	*(10* $^{-12}$ m)	MP(°C)
Rubidium	502	39
Cesium	540	23
Francium	600 (approx.)	25 (approx.)

Since these three elements also react with water, have a +1 valence and are similar in other ways, they too are included in the family. In order to show this similarity in properties these elements are placed in the same vertical column of the periodic table. In chemistry a family of elements is called a *group*. The family consisting of lithium, sodium, potassium, rubidium, francium and cesium is called Group I.

1 1 **H** 1.0
3 1 **Li** 6.9
11 1 **Na** 23.0
19 1 **K** 39.1
37 1 **Rb** 85.5
55 1 **Cs** 132.9
87 1 **Fr** 223

In most tables hydrogen is shown as part of Group I because of its +1 valence, but in many ways hydrogen is really an orphan and belongs to no family at all. Occasionally it is placed all by itself; occasionally it is listed in other families. You will notice that in our table we have listed it in two places. Of course, the other thing to note in our Group I family is that it is arranged in order of increasing atomic mass and atomic number as we move down the family.

Now, since calcium and magnesium are similar to the Group I elements in many ways, but different in other ways (such as valence number) let's find a place for them in the table. If we keep them close to Group I it can show that they are similar in some ways. Calcium must go below magnesium because it has the greater atomic mass and atomic number.

Magnesium has one more proton than sodium so let's make them neighbours. Calcium has one more proton than potassium so they should also be neighbours. Here is how things will look now.

1 ^{1} **H** 1.0	
3 ^{1} **Li** 6.9	
11 ^{1} **Na** 23.0	12 ^{2} **Mg** 24.3
19 ^{1} **K** 39.1	20 ^{2} **Ca** 40.1
37 ^{1} **Rb** 85.5	
55 ^{1} **Cs** 132.9	
87 ^{1} **Fr** 223	

We have two more families that we know of. One is the inert gas family. Arranged by increasing atomic mass we get the following group, which we will call Group VIII. Note the properties we have listed.

2 **He** 4.0
10 **Ne** 20.2
18 **Ar** 39.9
36 **Kr** 83.7
54 **Xe** 131.3
86 **Rn** 222

The Inert Gases

Table 5.5

	Specific Heat Capacity J/(kg·K)	*Boiling Point* °C	*Melting Point* °C	*Density* g/cm^3
He	1250	−269	−270	0.13
Ne		−246	−248	1.20
Ar	130	−185	−189	1.40
Kr		−152	−157	2.60
Xe		−108	−112	2.60
Rn		−60	−70	

Once again, note the regular change in properties down the table.

The third group of elements you investigated were the halogens. As before, you saw the similarities in chemical behaviour. If we arrange these elements in order of increasing atomic mass we get:

1 1 **H** 1.0
9 1 **F** 19.0
17 1 5 7 **Cl** 35.5
35 1 5 **Br** 79.9
53 1 5 **I** 126.9
85 1 **At** 210

Table 5.6

	Specific Heat Capacity J/(kg·K)	*Boiling Point* °C	*Melting Point* °C	*Density* g/cm³
F	750	−188	−219	1.50
Cl	500	−34	−100	1.56
Br	290	58	−7	3.12
I	210	183	113	4.94
At			estimated 300	

Notice that we have included the other two elements that have a −1 valence.

Before we continue we should point out one relatively minor thing. Hydrogen sometimes has a −1 valence. This occurs in the hydrides, for example, sodium hydride (NaH). Sometimes hydrogen is included at the top of the halogens. Incidentally, the halogens are called Group VII!

So far we have decided to list substances with similar chemical and physical properties in the same vertical column of the table. We also decided to place them in order of increasing atomic mass. To put these groups together we decided to move horizontally by increasing atomic number. If we now try to piece together the horizontal rows for the families we have looked at we arrive at this order.

1 1 **H** 1.0			1 1 **H** 1.0	2 **He** 4.0
3 1 **Li** 6.9			9 1 **F** 19.0	10 **Ne** 20.2
11 1 **Na** 23.0	12 2 **Mg** 24.3		17 1 5 7 **Cl** 35.5	18 **Ar** 39.9
19 1 **K** 39.1	20 2 **Ca** 40.1		35 1 5 **Br** 79.9	36 **Kr** 83.7
37 1 **Rb** 85.5			53 1 5 **I** 126.9	54 **Xe** 131.3
55 1 **Cs** 132.9			85 1 **At** 210	86 **Rn** 222

We have seen how properties change as we move down a group of the periodic table. It is now time to look at a horizontal row of the table. Because most of the substances are familiar to you, we will look at the third row of the table.

Across the Periodic Table

11	12	13	14	15	16	17	18
1	2	3	4	5 3	6 4 2	1 5 7	
Na	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
23.0	24.3	27.0	28.1	31.0	32.1	35.5	39.9

5.8 The Properties of Third-Row Elements

sodium
magnesium
aluminum
silicon
phosphorus
sulphur
chlorine
argon
litmus paper
test tubes
hydrochloric acid
sodium hydroxide
sodium oxide
magnesium oxide
aluminum oxide
silicon oxide
1 M phosphoric acid
1 M sulphuric acid

1. Carefully compare the properties of the elements found in the third row of the periodic table, inspecting the samples provided by your teacher. Look for similarities and differences particularly between elements that are adjacent in the table.
2. Carefully examine the oxides of these elements. Note similarities and differences.
3. Carefully examine the water solutions of these oxides. Use litmus paper to determine the acid-base properties of each solution.
4. Obtain 10 ml samples of each oxide solution in which the oxide shows evidence of being slightly soluble in water. Divide the sample into two equal parts. To one part add 6 M HCl a few drops at a time and shake the test tube. To the other part add 6 M NaOH and shake well.

5. In consultation with your teacher suggest other comparisons that might be made and, with prior approval, make these comparisons.

Questions

1. Write the formulae for the oxides of the third row elements in order, from left to right across the periodic table.
2. Write the formula for the product of the reaction between the oxide and water. Again go from left to right across the table. Which substances display both acidic and basic properties in that they are soluble in both acids and bases?
3. Make a general statement about the changes in properties of the elements of the third row of the periodic table as one moves from left to right across the table.
4. Suggest how the properties of elements in the second row of the table might change as one moves from left to right across the table.

As we see, there is a change in properties as we move horizontally across the table. Metallic properties give way to non-metallic properties. From sodium, which is very metallic, to chlorine and argon, which are definitely non-metallic, properties change in a fairly regular way. Remember too, that atomic number is increasing by one with each step to the left.

The same change takes place in the row that begins with potassium, a metal, and ends with bromine and krypton, non-metals. Lithium to fluorine follows the same pattern. If we look at the table horizontally we can almost spot the in-between element. This element represents the transition between the metallic left side of a row and the non-metallic right side of a row. The next experiment will give you a chance to work with one of these in-between elements.

5.9 Silicon

This experiment will give you a chance to devise your own tests and procedures.

Silicon, $Z = 14$, is between metallic aluminum and non-metallic phosphorus in the periodic table. It deserves special attention for a day.

silicon

14 4
Si
28.1

1. Your teacher will provide many different reagents for you to use in an effort to determine something about the chemical and physical properties of silicon. Use these materials to find out something about silicon. Be prepared to pool what you learn with other members of the class. Remember the safety rules. If there is doubt in your mind about trying anything, ask your teacher *before* you try it.

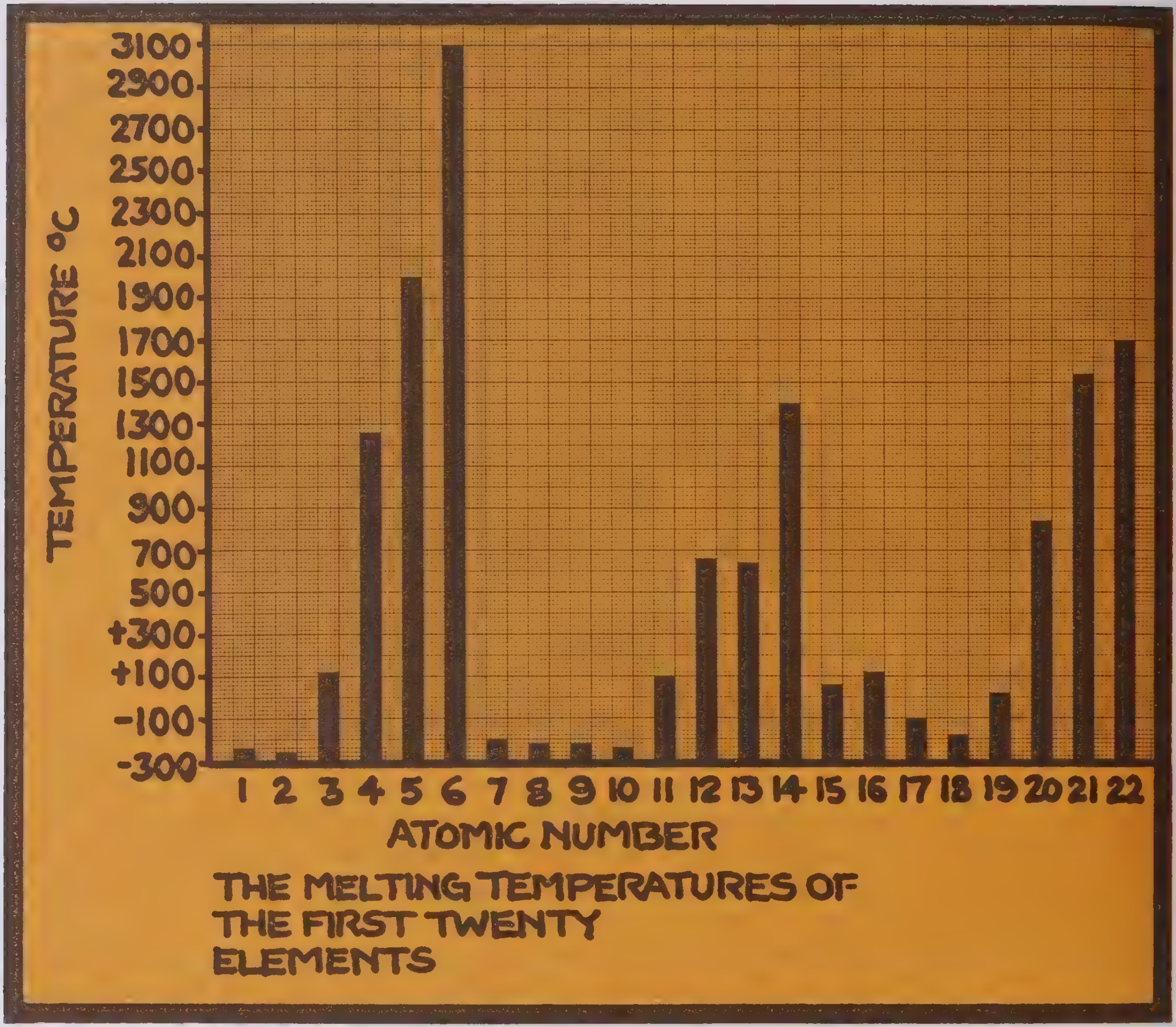
Questions

1. Is silicon a metal or a non-metal? Give reasons for your answer.
2. Silicon is often called a *metalloid*. What is the significance of that name?
3. Do you know of any other elements that show both metallic and non-metallic properties?

Graphical Analysis

We have seen that as the atomic numbers increase there is a repetition of certain physical and chemical properties. Let us now consider certain quantitative aspects of the elements in a slightly different form.

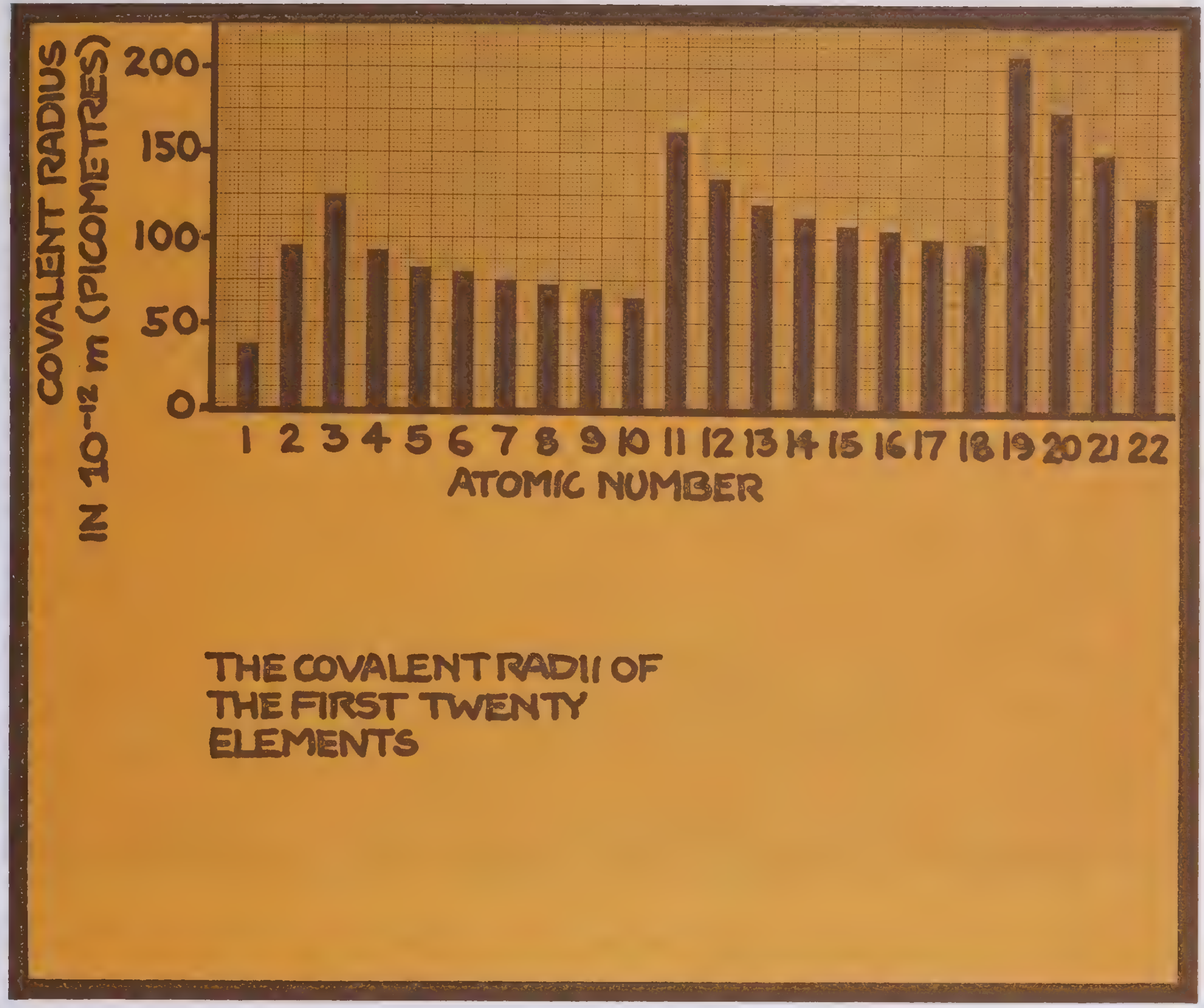
What follows are several graphs of certain properties or features of the elements plotted against atomic number. Study each graph carefully and then answer the questions. Remember that we are looking for similarities and differences.



Melting Point versus Atomic Number

Questions

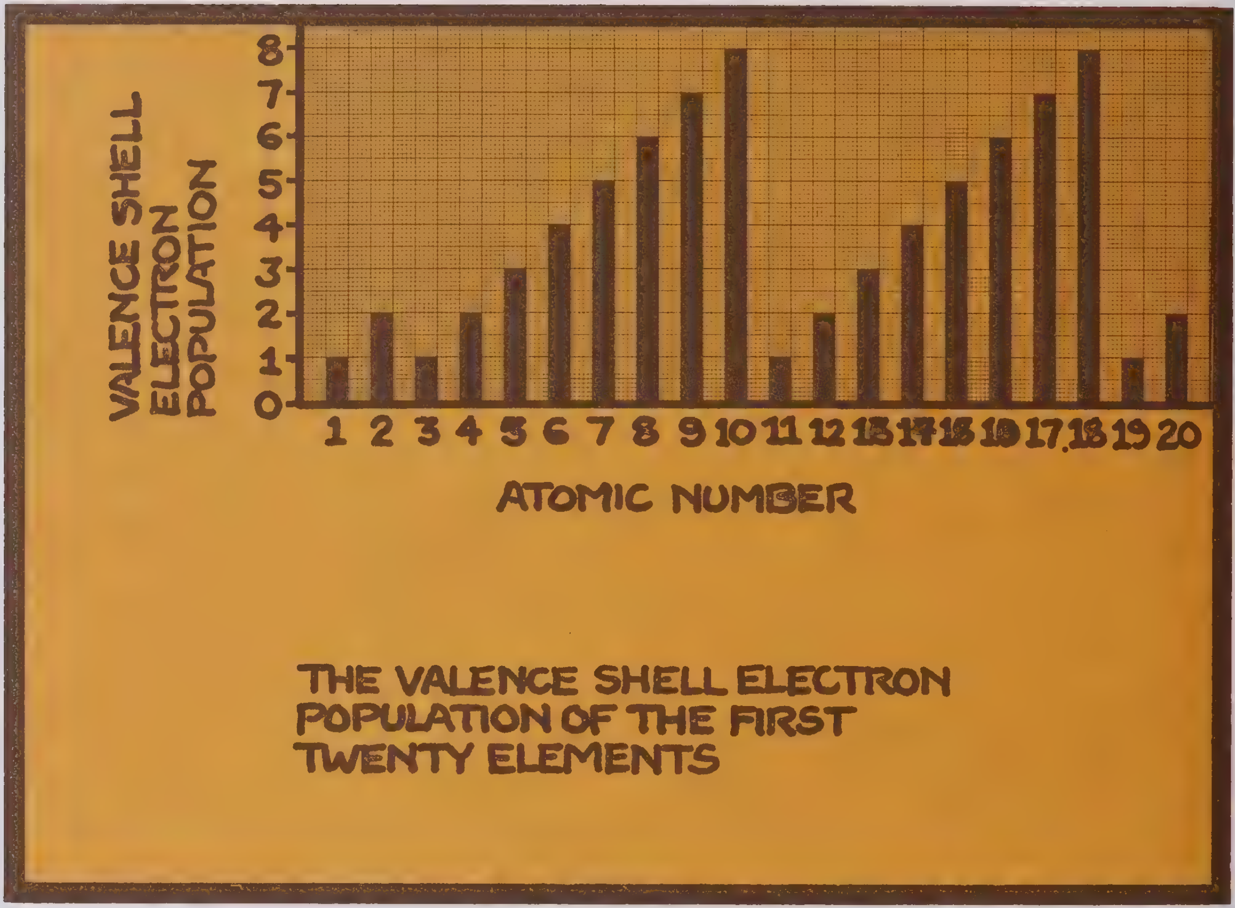
1. Describe the general shape of the graph.
2. Which elements represent the peaks?
3. Which elements represent the troughs?
4. For the first 20 elements only, list the elements between successive troughs.
5. How do your lists for Question 4 compare with the horizontal rows of the periodic table?
6. How do your answers for Questions 2 and 3 compare with the vertical columns of the periodic table?



Covalent Radius versus Atomic Number

Questions

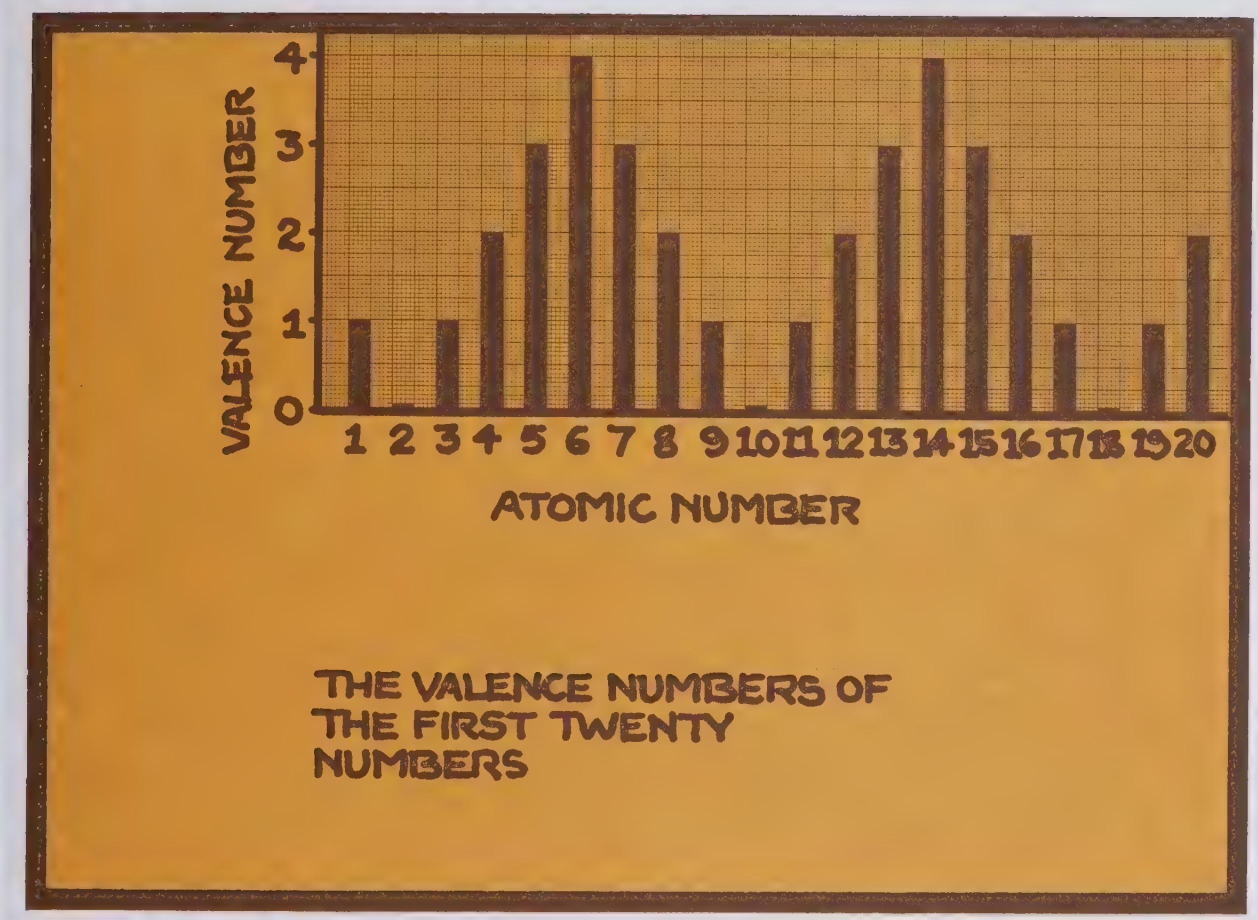
1. Describe the general shape of the graph.
2. Which elements represent the peaks?
3. Which elements represent the troughs?
4. How do your answers to Questions 2 and 3 compare with the vertical columns of the periodic table?
5. For the first 20 elements only, list the elements between successive peaks.
6. How do your lists for Question 5 compare with the horizontal rows of the periodic table?



Atomic Number versus Valence Shell Electron Population

Questions

1. Describe the general shape of the graph.
2. Which elements represent the peaks?
3. Which elements represent the troughs?
4. How do your answers to Questions 2 and 3 compare with the vertical columns of the periodic table?
5. For the first 20 elements only, list the elements between successive troughs.
6. How do your lists for Question 5 compare with the horizontal rows of the periodic table?

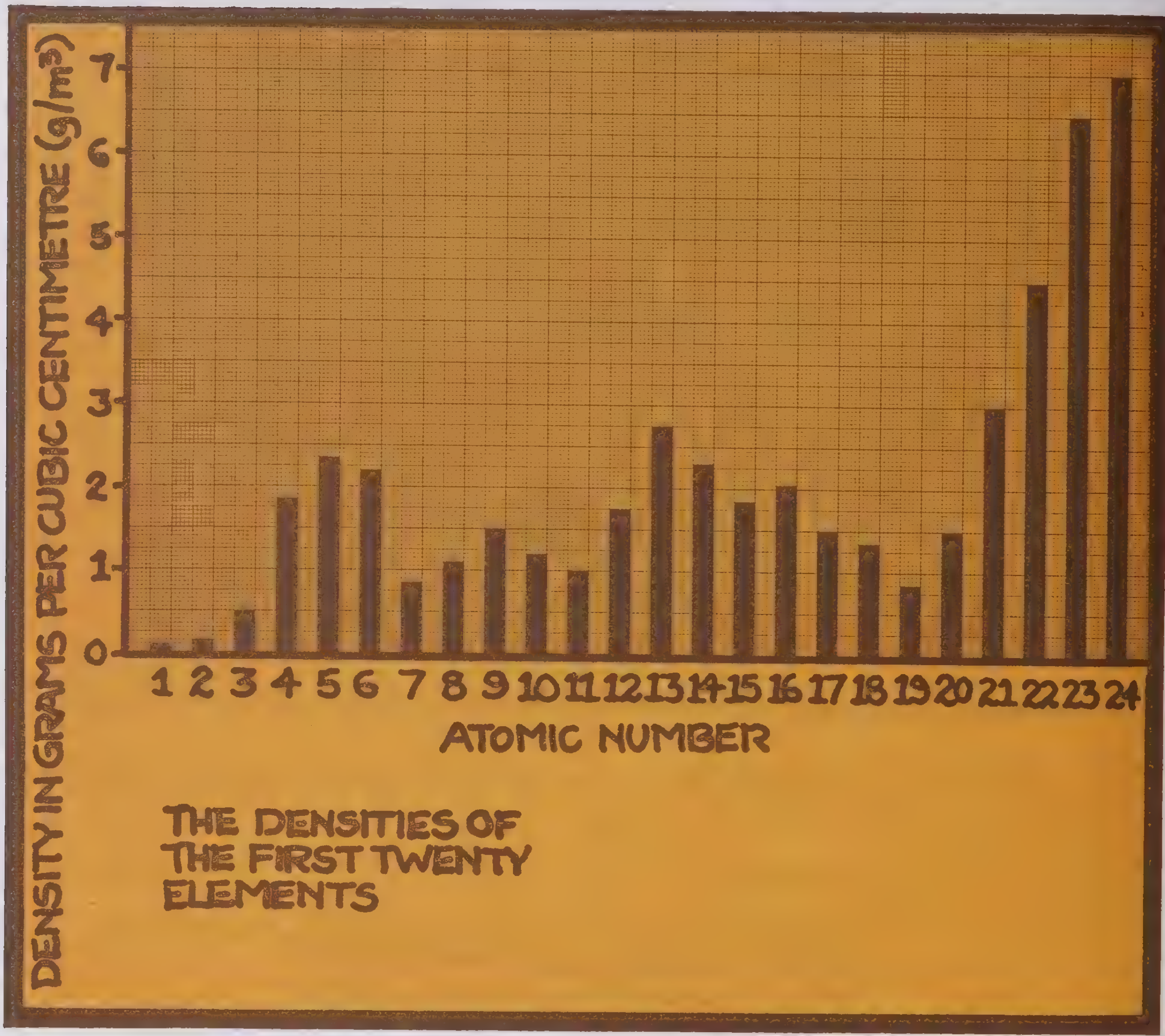


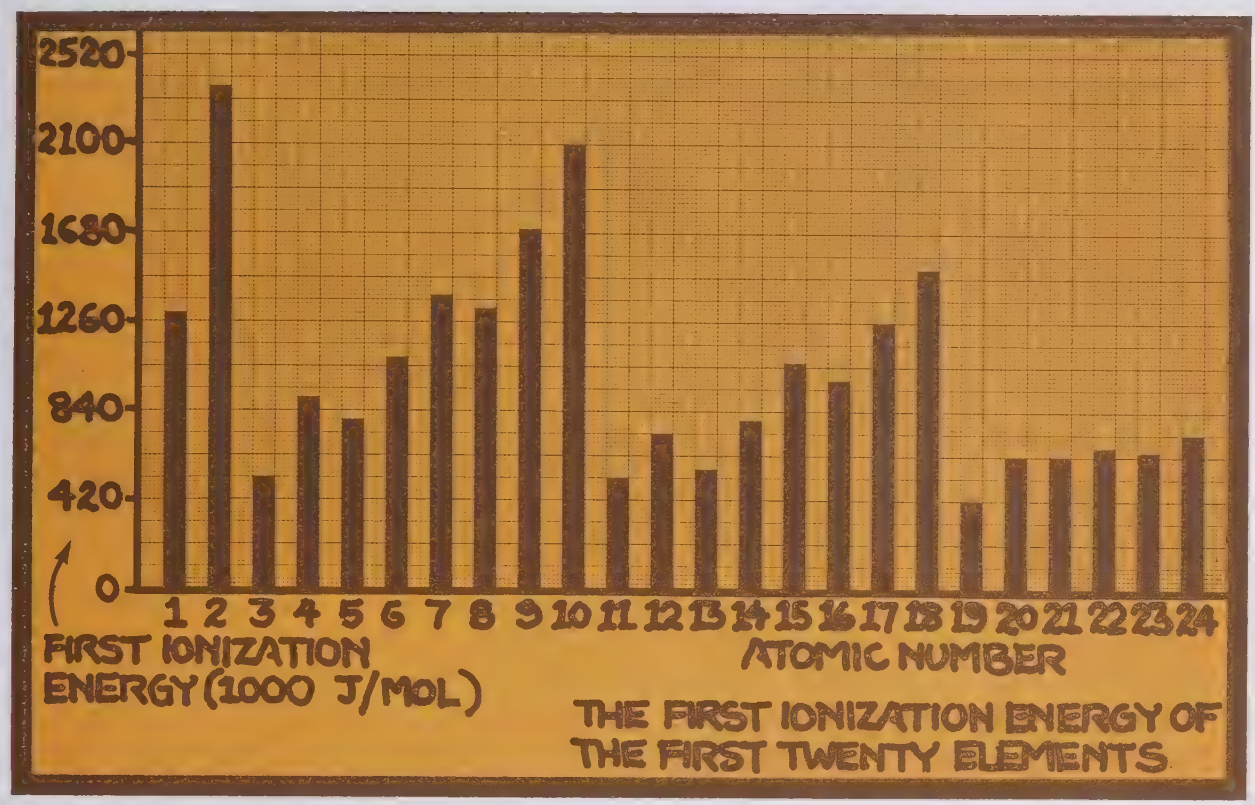
Atomic Number versus Valence Number (Absolute)

Questions

1. Describe the general shape of the graph.
2. Which elements represent the peaks?
3. Which elements represent the troughs?
4. How do your answers to Questions 2 and 3 compare with the vertical rows of the periodic table?
5. For the first 20 elements only, list the elements between successive troughs.
6. How do your lists for Question 5 compare with the horizontal rows of the periodic table?

Now we would like you to ask yourself some questions. The next two graphs represent properties of the elements. Write down the questions you would ask yourself about these graphs and then answer your own questions.





We hope the previous exercise has made its mark on you.

> "The properties of the chemical elements tend to repeat themselves in a regular manner."

This statement is known as the Periodic Law. It was revised by Henry Moseley after his work with x-rays to read as follows:

> "The properties of the elements are a periodic function of their atomic numbers."

The key to Moseley's Periodic Law is the single word *periodic*. That word is found in the title of the table that we know as the Periodic Table.

The word periodic means "repeating at regular intervals." Careful examination of the properties of the elements shows this repetition clearly. The key ideas for you to grasp about the periodic table are these:

1. The properties of the elements are a periodic function of their atomic number.
2. Elements in the same vertical column of the periodic table have similar chemical and physical properties.
3. Elements in the same horizontal row of the periodic table become more non-metallic as the atomic number increases.
4. Elements in the same vertical column of the periodic table have similar electron configurations in the outer occupied shell.
5. The atomic numbers of the elements increase by one with each move to the right on the periodic table. The atomic number represents the number of protons in the nucleus and the number of electrons in the neutral atom.

Linus C. Pauling

6. A vertical column of the periodic table is called a *group*. Each group is identified by a Roman numeral.
7. A horizontal row of the periodic table is called a *series* or *period*.
8. The metallic properties of the elements increase as one moves down the table.
9. The non-metallic properties increase as one moves to the right on the table.
10. The reactivity of non-metals increases as we move toward the upper right corner of the table (excluding the inert gas column of course). The reactivity of metals increases as we move toward the lower left corner of the table.

Problems

1. Without reference to your text, write the names, symbols and atomic numbers of the first 20 elements. (Hint: Try to duplicate the form of the periodic table.)
2. Classify each of the following elements as a metal, non-metal or metalloid:
   (a) strontium
   (b) astatine
   (c) boron
3. Which of the following pairs of elements would you expect to be more reactive?
   (a) magnesium or barium
   (b) tellurium or sulphur
   (c) cesium or lanthanum
4. Which element has the greatest tendency to gain electrons (become reduced)?
5. Which element has the greatest tendency to lose electrons (become oxidized)?

The Periodic Table

To do this final problem you will have to use all the knowledge you have gained in this course. It will not be easy but it may prove rewarding. Some preamble is necessary.

We would like you to rate all the elements on a scale of 10 to −10 on the basis of their metallic and non-metallic properties. We will score metals as negative values and non-metals as positive values. We have three reference points to help us.

First, we will rank all the inert gases as zero (0).

Second, since fluorine is the most non-metallic element we will rank it as +10.

Third, since francium is the most metallic element we will rank it as −10.

Now the only way to proceed is by comparing one element with another. We'll start with an example to get things going.

If fluorine is +10, then chlorine, bromine and iodine will be between +10 and 0 since all are non-metals. Also, chlorine will have a larger value than bromine; bromine will have a larger value than iodine. Astatine, the last member of the group, will have the lowest value. Let's make our comparisons, and rate them as follows:

fluorine	10
chlorine	8
bromine	6
iodine	4
astatine	2

Note that our values are purely arbitrary and represent *only* a comparison among these elements.

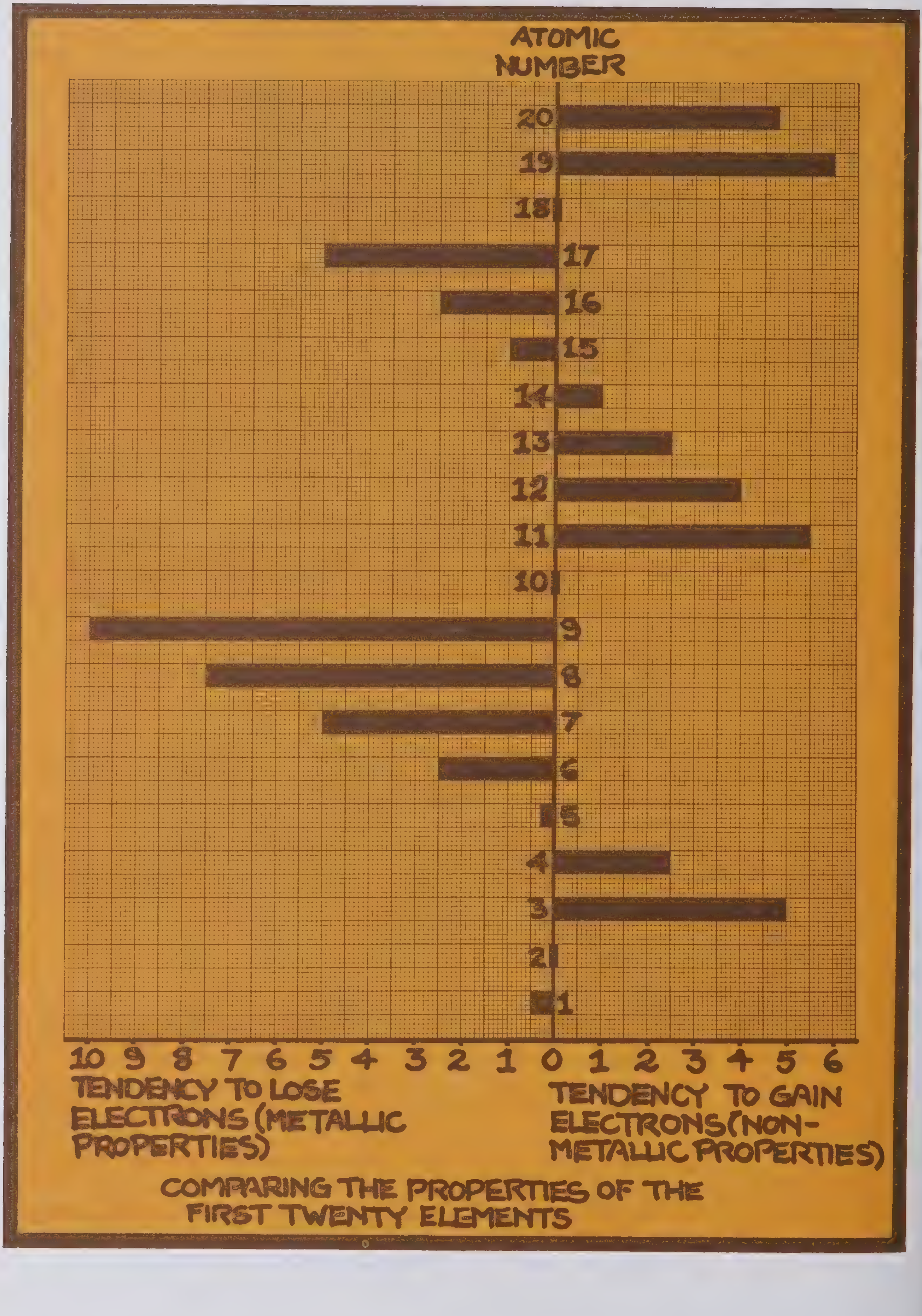
Similarly, on the other side of the table, francium is −10. The order must therefore be:

francium	10
cesium	
rubiduim	
potassium	
sodium	
lithium	?

Now we should remember that lithium, although at the bottom of this list, is certainly more metallic than some other elements. That is, lithium must not have *too* low a value. Shall we give lithium a value of −5? Can you fill in the values for the other members of Group I?

Now, continue until you have rated the first 20 elements and as many others as you are familiar with.

Finally, plot your values against atomic number on a graph. What sort of curve do you expect to get? Turn the page and you'll see our version.



COMPARING THE PROPERTIES OF THE FIRST TWENTY ELEMENTS

As you can see, this last graph shows the periodic change from metallic to non-metallic properties. Even if we argue over the values assigned we would still see the periodic nature of the data.

The most obvious point that arises from all this is that the periodic table can be divided into three regions:

1. metals
2. non-metals
3. inert gases

A comment about the inert gases first. They are not inert. Until 1963 they were thought to be inert. In that year, Neil Bartlett of the University of British Columbia was the first man to make an inert gas combine with another substance. Perhaps we should change their name to "noble gasses." This name might reflect their rather snobbish attitude toward the other elements.

The division between metals and non-metals is usually considered to be the one shown in the following periodic table. This compares rather favourably with our graph.

Knowledge of the basic patterns found in the table allow us to make predictions about elements that we are not familiar with. Many of these relationships carry over into the compounds that elements form. These relationships are best left to a second year of chemistry.

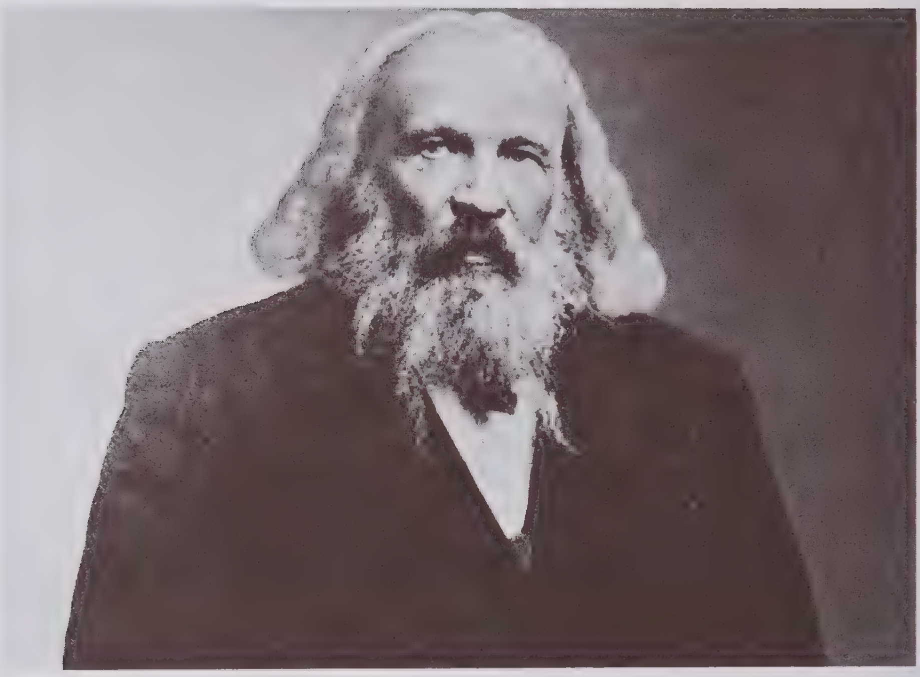
Last, but certainly not least, some knowledge of metallic and non-metallic properties gives us a clue to an element's tendency to form ionic or covalent bonds.

Elements that have a big difference in metallic and non-metallic properties tend to form ionic compounds. Elements that are not so widely separated on our graph tend to form covalent compounds.

The Periodic Table of the Elements as we know it was proposed by Dmitri Mendeleev in 1869. Mendeleev's table was a much improved and modified version of earlier tables proposed by John Newlands in 1864 and by Lothar Meyer in 1865.

Mendeleev arranged his table, as we did, in order of increasing atomic mass. Elements with similar chemical and physical properties were placed in vertical columns. The real brilliance of Mendeleev's work was that he recognized that several elements were missing. His table had several holes in it. These were spaces that he reserved for elements that he felt had not yet been discovered. Mendeleev predicted the properties of these missing elements quite accurately. He even indicated where these elements might be found. Again he was accurate. All these missing elements have now been found and so the present periodic table has no empty spaces.

Well, that's not exactly true. There are lots of empty spaces beyond element 105. If you would like to win a Nobel Prize in Chemistry maybe you could look for one of these still missing elements. More than one Nobel Prize has been given for the discovery of a new element.



Dmitri Mendeleev

Questions and Problems

1. From the descriptions given, decide whether the substance is a metal, non-metal or inert gas.
   (a) This element has a very low melting point.
   (b) This element is a very good conductor of heat and electricity.
   (c) This element forms ions with a positive charge.
   (d) This element forms an acidic oxide.
   (e) This element is not known to react with any other substance.
2. Given the following data on specific heat capacity, determine whether elements follow the Periodic Law.

Element	Specific heat J/(kg·K)
lithium	3300
beryllium	1900
boron	1300
carbon	700
sodium	1250
magnesium	1000
aluminum	900
silicon	700
phosphorus	880
sulphur	800
potassium	800
calcium	600

3. Prepare a list of properties for each of the major areas of the periodic table:
   (a) metals;
   (b) non-metals;
   (c) inert gases.
4. Indicate some of the properties that make it difficult to decide which section of the periodic table silicon belongs in.
5. The discovery of element number 109 has been reported in a scientific magazine. From your knowledge of the periodic table what predictions can you make about this new element?

Some Words You Should Know

element
compound
atomic mass
atomic number
metal
active metal
non-metal
halogen
activity
electronegative
electropositive
group
period (series)
row
column
metalloid
periodic
reduced
oxidized
inert gases
ionization energy
covalent radius

6

Destructive Distillation/Carbon/Petrochemicals/Hydrocarbons/Alcohols/Esters/Polymers

Carbon

Carbon is by far the most versatile chemical element that man has ever discovered. Carbon, unlike many of the other elements, can be found in the free form. By this we mean that carbon is found naturally as an uncombined element. In addition, carbon as a free element may exist in more than one form, as you will see later in this unit. Carbon combines with most of the other known elements to form binary compounds. The most unique property of carbon, however, is its ability to unite directly with other carbon atoms to form multi-carbon molecules that exist as chains, branched chains or ring-shaped molecules. Because of the unique ability of carbon to unite directly with other carbon atoms, several hundred thousand compounds of carbon are known. So broad is this branch of chemistry that it has become a separate area of study to many chemists. The branch of science concerned with carbon and its compounds is called organic chemistry.

Organic Chemistry

Organic chemistry is the study of carbon and its compounds wherever they are found. For the most part, however, organic chemistry centres on the study of carbon found in once-living things.

The word "organic" means existing as part of or derived from organisms. The early chemists felt that carbon-containing compounds could only be produced by living organisms, hence the name organic chemistry. In actual fact, however, the term "carbon chemistry" describes this branch of science better because now numerous compounds of carbon are produced synthetically. It was in 1828 when a man named Woehler first synthesized an organic compound called urea from known inorganic materials. However, since the word organic has been used in reference to carbon chemistry for so long, it still remains as part of the vocabulary.

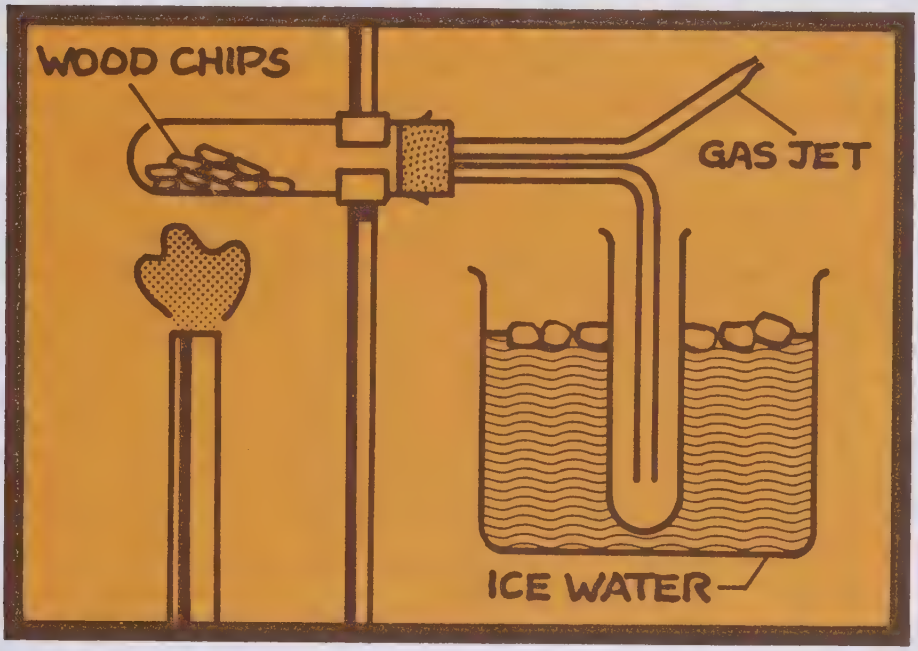
In the study of carbon chemistry we will first consider the element itself, where it is found and how it is used in the structure of various materials. Secondly, we will study some of the compounds of carbon and look at a systematic organization of a vast number of compounds.

6.1 Heating Substances out of Contact with Air

Record *detailed* observations as you proceed.

wood shavings
soft coal
paper
weed stalks
bone
6 test tubes
ice water

beaker
bunsen burner
retort stand and clamp
bent glass tubing
two-holed stopper
5 watch glasses



1. Arrange the apparatus as shown in the diagram.
2. Insert some wood shavings into the horizontal test tube. Reassemble the apparatus and heat the test tube with a moderate flame until there appears to be no further change.
3. During the heating bring a burning splint to the end of the gas jet.
4. After heating allow the contents to cool in each test tube. Carefully remove the contents of the horizontal test tube and put them in a watch glass for later examination.

Save each material for the next experiment.

5. Repeat the experiment with the other substances listed above.

You should now have five watch glasses each labelled indicating its origin.

Questions

1. What do you think was the most obvious change in each substance as a result of the heating? Describe this change.
2. Did each of the substances burn in the horizontal test tubes as you heated them? Did any of them burn? Upon what observations do you base your answer?
3. Did anything happen in the vertical test tube?
4. Did anything happen at the gas jet when a flame was brought near?
5. What one word might best describe what has happened to each of the original substances?
6. Compare the shape of each of the substances after heating to their original shape.
7. Do you think each of these substances had anything in common? Why do you think so?

6.2 Carbon in its Different Forms

powdered wood charcoal
coke
boneblack
lampblack
carbonblack
the products from Experiment 6.1
mortar and pestle
5 watch glasses

Warning
These substances are all very dirty.

1. Use a pestle to crush the remains in each watch glass from your previous experiment.
2. In additional watch glasses or on paper obtain samples of powdered wood charcoal, coke, boneblack, lampblack and carbonblack.
3. Compare the substances in the five watch glasses from your previous experiment to the different forms of carbon that you selected above. Comparison may be made on the basis of colour, feel, ability to mark paper or ability to burn.

Questions

1. What conclusion(s) can you draw about the five substances remaining at the end of Experiment 6.1?

2. What conclusion(s) can you draw about wood, paper, coal, bone and weed stalk?

Building Blocks of Nature

Each of the five substances in Experiment 6.1 was heated out of contact with air; hence none were able to burn. Heating done in this manner is called *destructive distillation*.

After the heating you probably found that each substance had become very brittle and black but that the basic shape of each was unchanged. Because of the properties of this black substance we are quite safe in assuming that it is carbon.

Because the basic shape of each of the original substances was unchanged and carbon is the only remaining substance it seems that the carbon atoms must form the framework upon which each of these substances is built. We could make a comparison here to a large building made of steel, brick, mortar, wood and glass. When such a structure is built, a framework of steel is made. This framework forms the basic shape of the building that will result. Added to this basic form are the brick, wood and glass. If such a building were strongly heated the wood and glass may disappear like the gases from coal or bone or paper when they were heated. The mortar and bricks may tumble to the ground like the liquids which resulted in Experiment 6.1. All that remains is the steel structure. It is very fragile now without the stability given it by the brick and mortar and wood, yet the shape remains.

In substances like those that we heated earlier, carbon atoms link together forming the framework of each. Atoms of other elements are linked to this carbon structure lending support and giving it stability but not contributing much to the form that the substance takes.

The carbon atom and its properties form the basis for the topic that we want to discuss. In abundance, carbon ranks eleventh by mass among the elements in the earth's crust, but, in importance it ranks much higher than eleventh, perhaps even first. The next logical step in our study of organic chemistry would be to find out just where carbon and carbon compounds are found.

6.3 Sources of Carbon

cotton
wool
leaves
flour
potato
carrot
cornstarch
sugar
meat
wax
lard
vegetable oil
crucibles and lids
bunsen burner
clay triangle
glass tube

To save time, it may be advisable for each lab group to work on only two or three substances, then pool their results.

1. The heating in each case will be done in crucibles with the lid on. The crucible should be supported by a clay triangle so that it may be heated directly.
2. Place a sample of one of the substances assigned to you into the crucible. Cover and heat with a hot bunsen flame for about 5 min.
3. Allow the crucible to cool, remove the lid and examine the contents.
4. Repeat the procedure with other substances.
5. Hold a glass tube in the flame of a candle. Remove it; then examine it.
6. Repeat Step 5 with a luminous (yellow) bunsen flame and with an alcohol burner in which kerosene is the fuel.

Questions

1. Were there any substances used that carbon did not seem to be a part? If so, list them.
2. Based upon your observations from this experiment and Experiment 6.1, list general areas where carbon appears to be found.
3. Explain the importance of carbon in substances such as bread, sugar and wood.

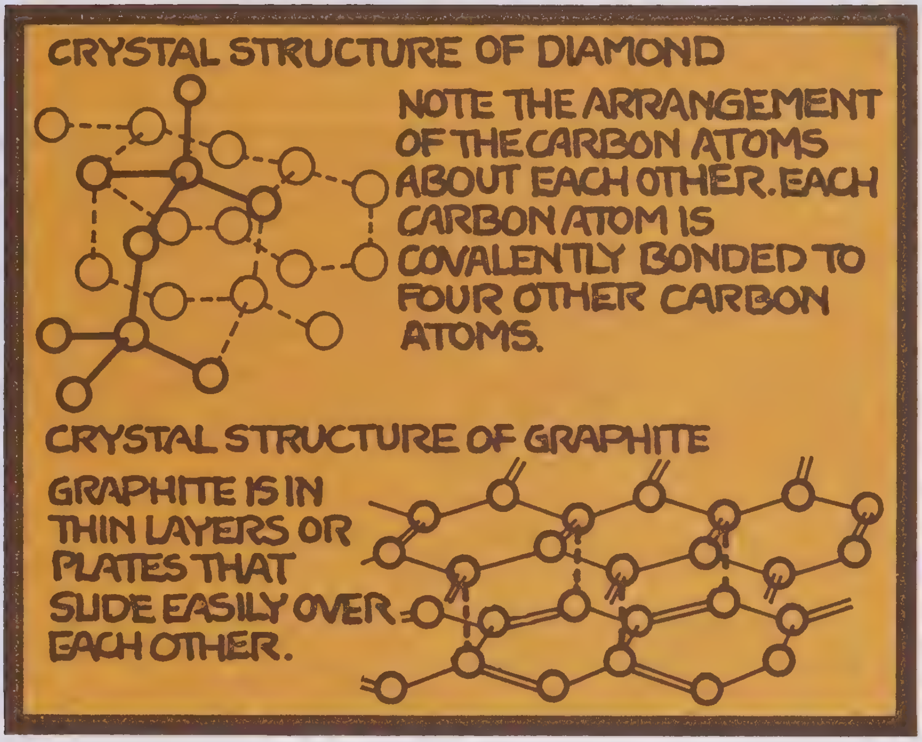
Forms of Carbon

If you were asked to describe the shape of the carbon particles that you have seen the word "shapeless" may come to mind. The particles in many instances are very tiny without any particular shape. The name given to carbon in this form is *amorphous*. This word originates from the Greek word meaning "without shape."

Other forms of carbon or allotropes of carbon do exist. One of the most beautiful sights in the world is that of a diamond, whether it be large or small. Diamond is a pure form of carbon. Crystal clear diamonds are cut for the gem trade. Less clear diamonds are used for drills or cutting tools because of the extreme hardness of the substance.

Diamond has a different set of properties from other forms of carbon because of the arrangement of the carbon atoms that build up the crystal. The lines and circles in the diagram represent a portion of a diamond. The relative positions of the carbon atoms are illustrated by the circles.

Graphite, a common lubricant for locks and gun bolts, is also a form of pure carbon. Graphite appears slippery when rubbed between the fingers because of its unique structure. Examination under a microscope reveals small flakes or plates that slide easily over each other. Notice the difference between diamond and graphite by comparing the two line and circle diagrams.



CRYSTAL STRUCTURE OF DIAMOND

NOTE THE ARRANGEMENT OF THE CARBON ATOMS ABOUT EACH OTHER. EACH CARBON ATOM IS COVALENTLY BONDED TO FOUR OTHER CARBON ATOMS.

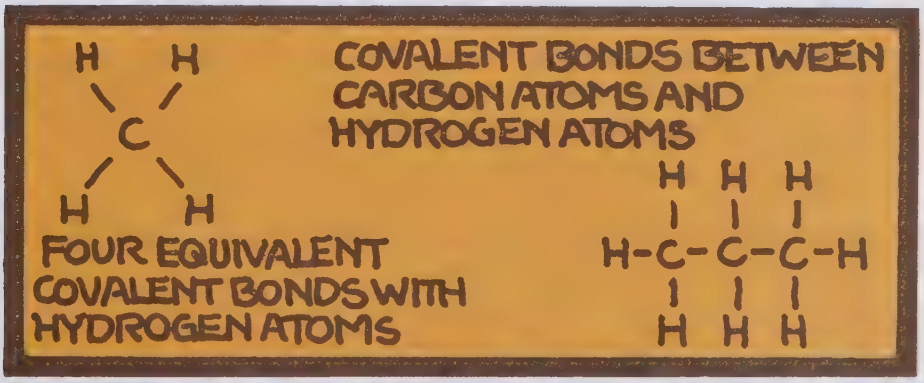
CRYSTAL STRUCTURE OF GRAPHITE

GRAPHITE IS IN THIN LAYERS OR PLATES THAT SLIDE EASILY OVER EACH OTHER.

The more common forms of carbon that we will examine in our study of the element carbon will fall mainly into the amorphous forms. Let us look more closely at some amorphous carbon and some of its properties.

The Carbon Atom

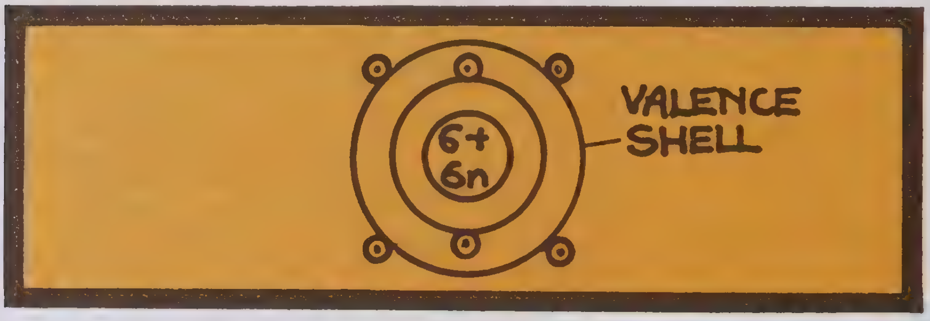
The carbon atom is the most versatile element known to man. It is found in Group IV A of the periodic table, exactly mid-way between helium and neon, two inert gases. The carbon atom has an atomic number of 6, hence each atom has six extra nuclear electrons with four of these found in the valence shell.



COVALENT BONDS BETWEEN CARBON ATOMS AND HYDROGEN ATOMS

FOUR EQUIVALENT COVALENT BONDS WITH HYDROGEN ATOMS

The four valence electrons are capable of forming four equivalent covalent bonds. These bonds may be with atoms of another element or with carbon atoms and atoms of another element.



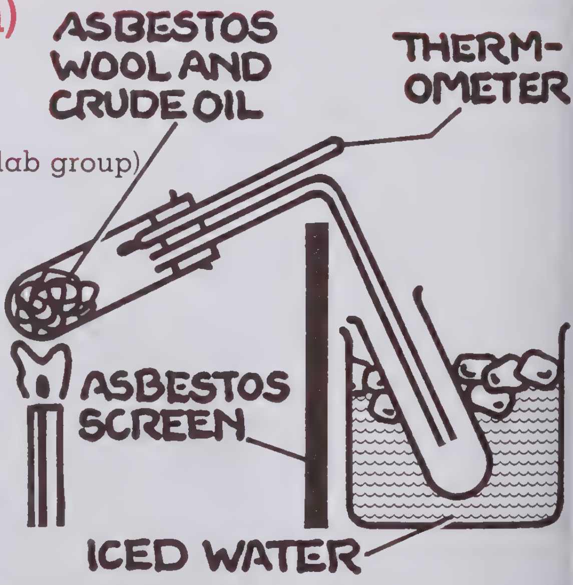
Because carbon atoms can combine covalently with other carbon atoms, long chainlike molecules may result. These molecules may be in chains or in chains with branches from them. Each different chain forms a new molecule with a new set of properties. Molecules with more than 50 carbon atoms are presently known. It should be apparent to you that there are an almost unlimited number of ways of combining carbon atoms, leading to literally thousands of carbon compounds.

At the start of this unit, we found that carbon atoms are the basic building blocks for once-living things. We want now to turn our attention from the element to compounds of carbon and their sources. We will begin by distilling some crude oil.

6.4

Carbon Compounds from Crude Oil (Demonstration)

asbestos wool
crude oil (about 4 ml per lab group)
6 test tubes
two-holed stopper
beaker
ice water
thermometer
bent glass tubing
bunsen burner
retort stand and clamp
crucible
splint



1. Set up the apparatus as shown in the diagram.
2. Put a loose asbestos plug in the bottom of the horizontal test tube (check diagram).

3. Add about 4 ml of the crude oil on top of the asbestos plug. This should be done without getting the oil on the sides of the test tube.
4. Reassemble the apparatus and lower the vertical test tube into the beaker of cold water.

A fraction is the portion collected over each temperature range.

5. The heating should be done slowly and gently. Fractions may be collected over the following temperature ranges:
   (a) from initial temperature to 75°C
   (b) 75°C to 125°C
   (c) 125°C to 165°C
   (d) 165°C to about 210-220°C.

Change the vertical test tube after collecting each fraction.

6. Describe each of the four fractions. In your description include such things as colour, odour and how readily each pours.
7. Remove a small sample of the distillate from the first test tube and place it in a crucible. Light the sample with a burning splint. Describe the flame. Repeat with other fractions.

If a sample does not ignite readily put a small amount of asbestos wool in the crucible for a wick.

8. Mix all that remains in the test tubes back together again. Compare this mixture to the original crude oil.

Questions

1. Give ways by which you know that each fraction was different from each other and from the crude oil.
2. What was the single most distinctive feature of the flame produced as each fraction burned?
3. Recall the candle flame or a bunsen flame with a poor air supply. What caused these flames to have their distinctive yellow colours?
4. What conclusion can you draw about the constituents of each of the fractions?
5. From the odour or the colour of each fraction can you name a common substance that any one of them resembles?

Petroleum or crude oil, is by far the greatest source of liquid fuel in the world today. Petroleum is a black, watery liquid consisting of a complex mixture of compounds of hydrogen and carbon. Such compounds are referred to as *hydrocarbons*. The mixture consists of numerous liquid hydrocarbons but, when petroleum is carefully distilled, various gaseous as well as solid hydrocarbons are collected.

Many useful products are obtained from petroleum in the distillation process. Among the more valuable are gasoline, furnace fuel oil and lubricating oils.

Oil refinery at Calgary, Alberta. Crude oil is pumped into towers such as these where it is separated by distillation into many valuable products such as gas, gasoline, jet fuel, diesel fuel, heating oils and lubricants. *Courtesy of Imperial Oil Ltd.*

Compounds of Carbon with Hydrogen

The simplest class of carbon compounds are those that contain only carbon and hydrogen. Compounds of this sort are called *hydrocarbons*. When all the carbon atoms in a molecule are joined by single covalent bonds the resulting molecules are termed alkanes or paraffin hydrocarbons. The two synonymous names both reveal something to us about the compounds. *Paraffin* is a word of Latin origin. *Parum* means little and *affinis* means affinity. A hydrocarbon with single covalent bonds between carbon atoms has little affinity to enter reactions with other substances.

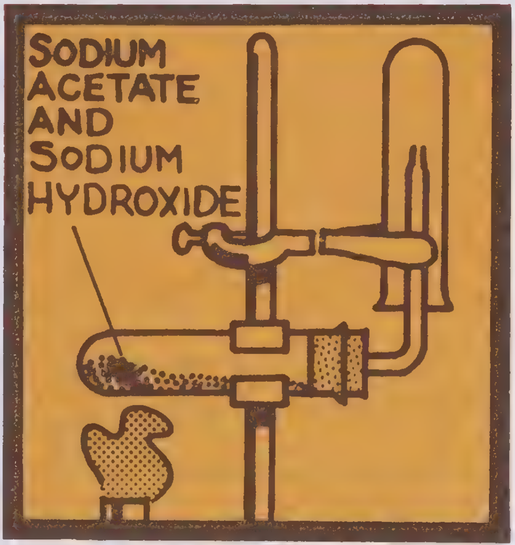
The name *alkane* is applied to this class of compound in a systematic method of describing the structure of these molecules. The suffix -ane indicates to us molecules with carbon atoms each joined by single covalent bonds. Specific simple alkanes are:

methane — CH_4
ethane — C_2H_6
propane — C_3H_8
butane — C_4H_{10}
pentane — C_5H_{12}
hexane — C_6H_{14}
heptane — C_7H_{16}
octane — C_8H_{18}

Each successive compound in the alkane series increases by one carbon atom and two hydrogen atoms. Compounds of carbon that fall into the alkane or paraffin series are termed *saturated* hydrocarbons. Each carbon atom has the capacity for four covalent bonds. If each of these bonds are single covalent bonds and if each bond is to another carbon atom or to a hydrogen atom, then the molecule is a saturated hydrocarbon.

6.5 To Prepare and Study the Properties of Methane

sodium acetate
sodium hydroxide (soda-lime)
limewater
2 test tubes
bunsen burner
retort stand and clamps
one-holed stopper
bent glass tubing
gas jar



Reaction: $CH_3COONa + NaOH \longrightarrow Na_2CO_3 + CH_4(g)$ (methane)

1. Prepare a mixture of equal portions of sodium acetate and sodium hydroxide. Make certain that the two are well mixed. Place about 2 cm of the mixture in a test tube then set it up in the arrangement shown in the diagram.
2. Tap the tube to make certain that the mixture is spread along the length of the test tube (see diagram).
3. Begin the heating by moving the bunsen flame along the test tube. After 1 to 2 min lower a second test tube over the open end of the L-tube.

Leave the tube in this position for about 1 min then remove it and take it at least a metre from the bunsen flame. Bring a burning splint to the mouth of the inverted test tube.

Warning

Being too eager at this point can be dangerous. Follow the instructions in Step 4 very carefully.

4. Repeat this procedure until the gas collected ignites and burns quietly. Now, light the gas at the end of the L-tube. Describe the flame produced with respect to structure and luminosity.
5. Hold a clean dry gas jar over the flame. Describe any deposit. After about 30 s turn the jar upright and add a few millilitres of limewater. Cover and shake the bottle. Describe any changes in the limewater.

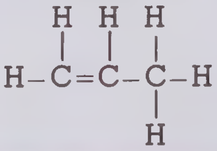
Questions

1. What colour was the flame at the end of the outlet tube?
2. What has been the cause of this colour flame on previous occasions?
3. What colour was the deposit that formed on the inverted gas jar?
4. What was the deposit on the gas jar?
5. What happened to the limewater when shaken with the gas produced by the flame?
6. What substances are produced when methane burns?
7. Attempt to write an equation for the burning of methane in air.

Alkenes

In many instances, however, two carbon atoms in a hydrocarbon are joined by a double covalent bond. These compounds, as we will find, have properties somewhat different from the alkanes. This means that a new series of hydrocarbons is formed by the insertion of a double bond between two of the carbon atoms in a molecule. To indicate this by name the series is referred to as the *alkenes* and each member of the series has a name that ends in the suffix -ene. Specific members of this series are:

ethene C_2H_4
propene C_3H_6
butene C_4H_8
pentene C_5H_{10}



$$\mathrm{H-\overset{\overset{H}{|}}{C}=\overset{\overset{H}{|}}{C}-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-H}$$

Graphic formula for *propene*. Note here one double bond and four covalent bonds for each carbon atom.

6.6

Alkanes and Alkenes Compared

suitable alkanes and alkenes such as:
cyclo-hexane
cyclo-hexene
benzene
toluene
other reagents:

0.01 M $KMnO_4$ (acidified)
bromine water
concentrated H_2SO_4
10 test tubes

Warning

Extreme danger may be encountered with bromine water and concentrated sulphuric acid. If these are to be used by students every precaution must be taken and directions must be followed precisely. Avoid contact with the skin.

1. Place about 3 ml of the $KMnO_4$ solution into the bottom of four clean test tubes. Add about three drops of an alkane to one of the test tubes, then shake it gently. Repeat with each of the other alkanes or alkenes provided.
2. Follow a similar procedure to Step 1 above only use the bromine water instead of the $KMnO_4$ solution.
3. Demonstration by teacher.
   Place three drops of an alkane in one test tube and the same amount of an alkene in a second test tube. Add one drop of concentrated H_2SO_4 to each test tube.

Questions

1. Which of the hydrocarbons tested were reactive with $KMnO_4$? Br_2 in H_2O? H_2SO_4 (concentrated)?
2. Examine a ball and stick model of each of the hydrocarbons used in the experiment. Which of the hydrocarbons used contain double bonds?
3. Does there seem to be any relationship between the reactivity of the hydrocarbons and the type of bonds found within the molecule? Make any generalization that seems apparent.

Triple Bonds

To carry our discussion of hydrocarbons one step further it should be noted that there are instances where a triple covalent bond exists between two carbon atoms in a carbon chain. This means, of course, a third series of hydrocarbons and a new name to indicate that the molecules of this series are somewhat different from the molecules of the alkene series. The name given to this new series is *alkyne*. The suffix -yne indicates to us a triple bond between two carbon atoms in a carbon chain. The simplest and most common member of this series is ethyne or better known to you and me as acetylene.

Graphic formula for *acetylene*. $H-C\equiv C-H$

A triple bond exists between the two carbon atoms. Note also the four bonds associated with each carbon atom.

6.7

Preparation and Properties of Acetylene

calcium carbide
limewater
red and blue litmus
400 ml beaker
small beaker
4 test tubes
litmus paper

Warning

Do not handle acetone near an open flame. Acetone must only be used in a well-ventilated area.

1. Half fill a 400 ml beaker with water and invert three 18 mm by 150 mm test tubes filled with water into the beaker.
2. Drop a lump of calcium carbide into the water and place each test tube in turn over the lump of material. Allow the first test tube to be completely filled with gas, the second test tube half full and the third about 1/10 full.
3. Place your thumb over the end of the first test tube and remove it from the beaker. Ignite the contents of the tube over the bunsen flame. Describe the flame and any deposit formed on the test tube. When combustion is complete turn the test tube upright, add a few millilitres of limewater and shake the tube.
4. Repeat Step 3 with the other two test tubes. Allow the water to run out of the test tube before placing your thumb over the end.
5. Test the solution in the beaker from Step 1 with red and blue litmus.
6. Filter a portion of the solution that is in the beaker then blow your breath through the filtrate for a few minutes.

Questions

1. What evidence did you gather that would indicate that a reaction had taken place between the calcium carbide and the water?
2. What substance did you collect in each test tube? What was different about each test tube's contents?
3. Describe the manner of burning in each of the first three test tubes.
4. What deposit did you find at the mouth of the first tube after burning the gas that was in it?
5. Did you find this same deposit at the mouth of the second and third test tubes?

6. How would you describe the combustion in each of the three test tubes (complete or incomplete)?
7. What factor seems to control the completeness of combustion of acetylene?
8. How does a welder ensure complete combustion of acetylene in order to obtain a good hot flame?
9. What product of combustion is indicated by the limewater test?
10. What change occurred in the solution from Step 6 by filtering?
11. What happened to the filtrate when you exhaled through it? What does this change tell you about the solution from the beaker in which the calcium carbide was placed?
12. The gas produced in Step 2 is acetylene. Write a balanced chemical equation for the reaction between calcium carbide and water.
13. Write balanced chemical equations for the incomplete and the complete combustion of acetylene.

Alkynes

Notice that in each molecule every carbon atom has four covalent bonds.

Both of these reactions are used as tests for the presence of a double or triple bond.

Of the alkyne series, acetylene is by far the most common. Information on alkynes beyond octyne (eight carbon atoms) is somewhat limited.

Acetylene is a colourless gas at room temperature with a rather distinctive odour. One of the most important physical properties of the gas is its extreme solubility in acetone. This is about 25 volumes to 1 at 100 kPa pressure rising to near 300 volumes to 1 at 1200 kPa pressure. Because of this property acetylene is delivered to welding companies under pressure but in acetone solution.

Alkynes undergo chemical reactions much like the alkenes. They add bromine or react with permanganate just as we found alkenes to do.

Alkynes like acetylene undergo complete combustion in air to produce carbon dioxide and water:

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

In a limited supply of air the combustion becomes incomplete, yielding large quantities of carbon:

$$2C_2H_2 + 2O_2 \longrightarrow CO_2 + 2H_2O + 3C$$

Summary

The alkanes, when pure, are colourless and almost completely without odour. The alkanes are almost completely insoluble in water. They all have densities less than that of water and methane has a density less than air. The lower molecular mass alkanes are gases up to butane; the intermediate molecular mass alkanes are liquids and the remainder, those of high molecular mass (with about 17 carbons or more) are solids.

As the name of the series (paraffin) implies, the alkanes are relatively unreactive. The single most important reaction that alkanes undergo is oxidation. With a sufficient oxygen supply, members of this series of hydrocarbons undergo complete combustion, yielding carbon dioxide and water:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

Under the less ideal conditions alkanes undergo incomplete combustion, yielding products such as water, carbon monoxide, carbon dioxide and carbon.

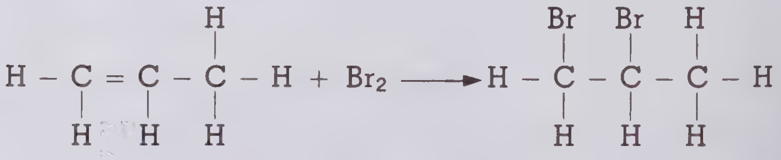
Alkenes resemble alkanes very much when their physical properties are compared. They also are colourless and almost odourless with densities less than water. Again, alkenes up to and including butene are gases while others are liquids or solids.

The reactivity of alkenes, however, is somewhat different from that of the alkanes. This difference we discovered in Experiment 6.6, when we compared members of the two series. Alkenes have the ability to decolourize potassium permanganate or bromine water.

The reaction of permanganate with an alkene is termed *oxidative addition*. The alkene is oxidized by the permanganate as shown in the reaction below:

$$3C_2H_4 + 2MnO_4^{-} + 4H_2O \longrightarrow 3CH_2OHCH_2OH + 2MnO_2 + 2OH^{-}$$

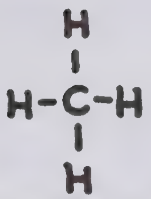
The decolourization of bromine by an alkene is an example of a straight addition to the alkene molecules. One of the double bonds opens, adding a bromine atom to adjacent carbon atoms. A typical reaction is given in the diagram:

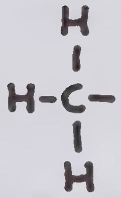


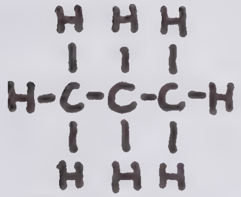
$$\text{H}-\underset{\text{H}}{\overset{}{\text{C}}}=\underset{\text{H}}{\text{C}}-\underset{\text{H}}{\overset{\text{H}}{\text{C}}}-\text{H}+\text{Br}_2\longrightarrow\text{H}-\underset{\text{H}}{\overset{\text{Br}}{\text{C}}}-\underset{\text{H}}{\overset{\text{Br}}{\text{C}}}-\underset{\text{H}}{\overset{\text{H}}{\text{C}}}-\text{H}$$

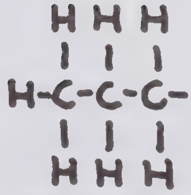
Branched Molecules

When a derivative is made from a simple hydrocarbon molecule, the hydrocarbon part of the molecule is called an *alkyl group* or the *alkyl* part of the molecule. For example, if a hydrogen atom is removed from a methane molecule and the remaining part of the molecule becomes attached to a layer molecule then the methane part of this molecule is an alkyl group called the *methyl* group.


METHANE

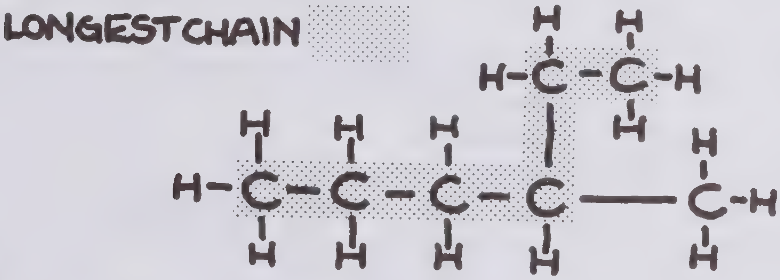

METHYL GROUP


PROPANE


PROPYL GROUP

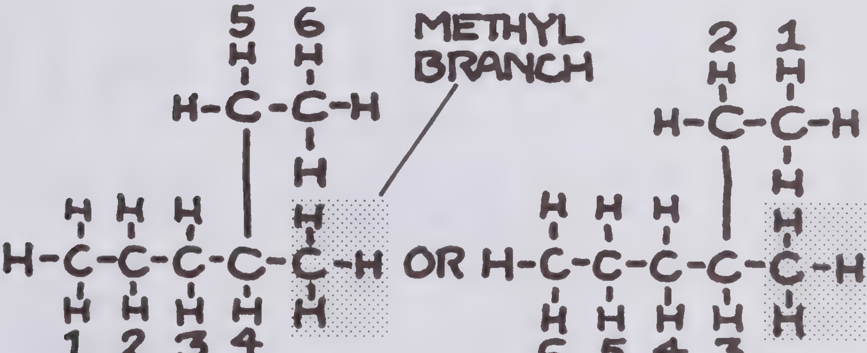
In naming a molecule that has one or more alkyl groups as part of it, the following rules are followed:

1. Count the number of carbon atoms in the *longest* continuous chain of atoms. This will form the basis for the name. For example:



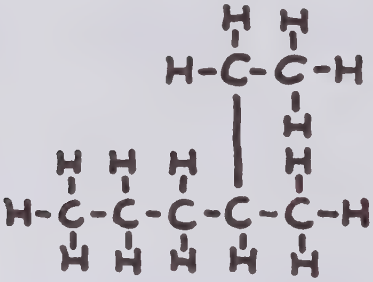
The molecule shown above contains six carbon atoms in the longest continuous chain. This means that the prefix *hex* appears in the name.

2. Examine the molecule for the type of bonding (single, double or triple) between carbon atoms. This will determine the suffix for the name. In the example above, all carbon-to-carbon bonds are single so the suffix in the name is *ane*.
3. Examine the molecule for branches. Name them. In the molecule shown, only one branch appears. This branch is a *methyl* group (one carbon).
4. Number the position of the branch on the main chain. Number the carbon atoms in the main chain to get the smallest possible number (or total of numbers in a multi-branched molecule).



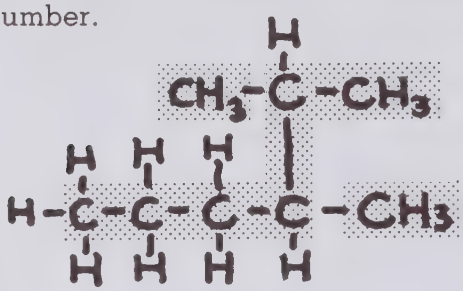
In the example selected we would number the carbon atoms in the main chain from the right end of the molecule. The methyl branch is on the number three carbon.

5. Put each of the pieces together beginning with the number of the *smallest branch* if there is more than one.



3-METHYL HEXANE

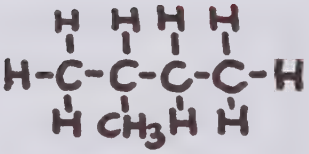
Notice that a dash separates a number from a word. If more than one number appears in succession a comma separates each number.



2,3-DIMETHYL HEXANE

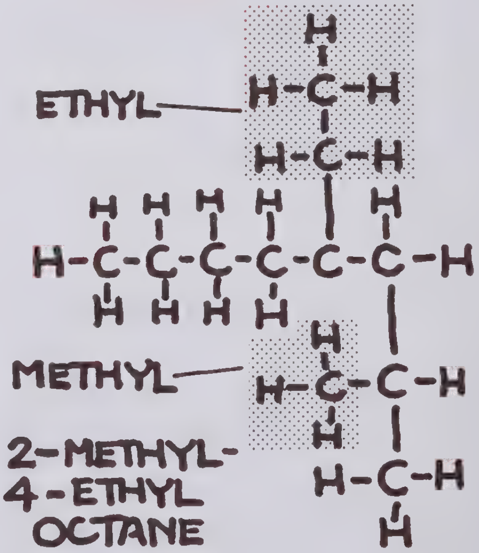
In the example above the branches are again numbered from the right to give the smallest total set of numbers. The prefix *di* is used to denote two methyl groups as branches.

Example 1



2-METHYL BUTANE

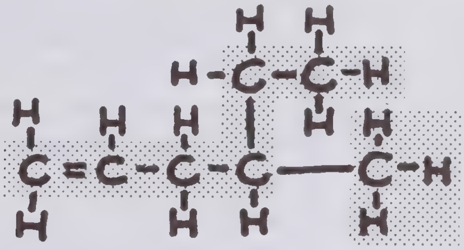
Example 2



2-METHYL-4-ETHYL OCTANE

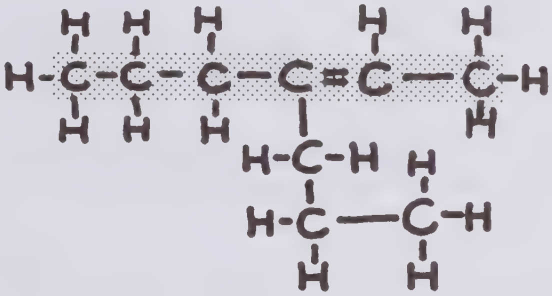
Branched Alkenes or Alkynes

If an alkene or alkyne has a branch from the main chain, the same rules for naming the molecule are used as described earlier but with two small changes. The position numbering is done from the end of the molecule that gives the double or triple bond the *smallest* number.



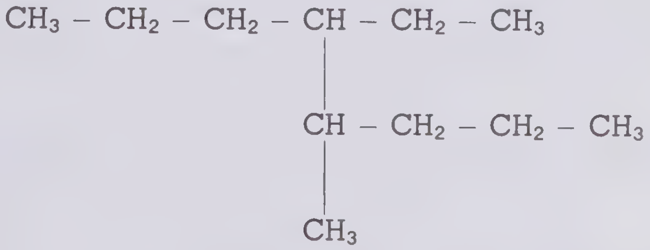
The name for this compound is: 4-methyl-1-hexene *not* 3-methyl-5-hexene

Secondly, the longest carbon chain must contain the double or triple bond, whichever the case may be. For example, in the molecule below, the longest chain that forms the basis for the name contains only six carbon atoms since the double bond *must* appear in this chain.



Problems

The following questions are about this formula:



$$\begin{array}{llllll}
CH_3 - & CH_2 - & CH_2 - & CH & - CH_2 - & CH_3\\
 & & & | & & \\
 & & & CH & - CH_2 - & CH_2 - CH_3\\
 & & & | & & \\
 & & & CH_3 & &
\end{array}$$

1. (a) How many carbon atoms are in the longest continuous chain?

(b) Since the compound contains only carbon and hydrogen atoms all with single bonds, to what group of organic compounds does it belong?

(c) Copy the formula into your notebook and draw a line through the longest carbon chain.

(d) How many branches are there? Name them.

(e) Will the numbering of the branches be done from the left or from the right? Explain.

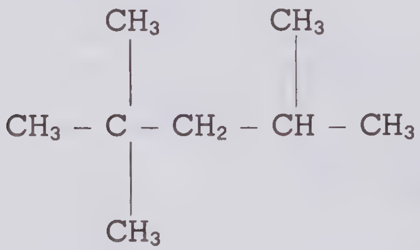
(f) Number the position of each branch.

(g) Name the branches from the smallest to the largest. (If a branch appears more than once, use a prefix to indicate this number.)

(h) Name the longest carbon chain and give it the proper ending.

(i) Write a complete name for the compound using your answers to Questions f, g and h.

2. (a) How many carbon atoms has the carbon chain forming the basis for the name of this compound?

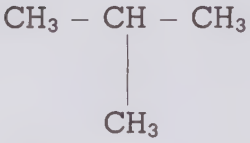


$$\begin{array}{ccccccccc} & & CH_3 & & & & CH_3 & & \\ & & | & & & & | & & \\ CH_3 & - & C & - & CH_2 & - & CH & - & CH_3 \\ & & | & & & & & & \\ & & CH_3 & & & & & & \end{array}$$

(b) Name the compound.

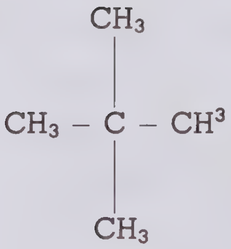
3. Name the following:

(a)

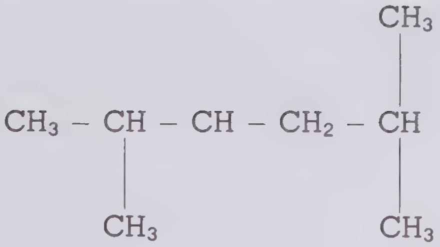


$$\begin{array}{ccccc} CH_3 & - & CH & - & CH_3 \\ & & | & & \\ & & CH_3 & & \end{array}$$

(b)



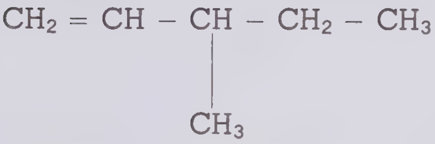
$$\begin{array}{ccccc} & & CH_3 & & \\ & & | & & \\ CH_3 & - & C & - & CH^3 \\ & & | & & \\ & & CH_3 & & \end{array}$$

(c) 

$$CH_3 - \underset{\displaystyle CH_3}{\underset{|}{CH}} - CH - CH_2 - \overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{CH}}}}$$

(d) $CH_3 - CH = CH - CH_3$

(e) $CH_2 = CH - CH_2 - CH_2 - CH_3$

(f) 

$$CH_2 = CH - \underset{\displaystyle CH_3}{\underset{|}{CH}} - CH_2 - CH_3$$

4. Draw structural formulae for:
   - (a) 2-methyl butane;
   - (b) 2-2-dimethyl hexane;
   - (c) 2-methyl-3-ethyl heptane;
   - (d) 4-methyl-2-pentene;
   - (e) 3-methyl-3-heptene;
   - (f) 7-methyl-2-octene.

Functional Groups and Chemical Behaviour

The most distinctive chemical properties of alkanes was their ease of oxidation to carbon dioxide. If we break this oxidation down into a series of steps, each step adding one additional oxygen atom, we would find that a number of intermediate compounds result before we actually get to carbon dioxide.

For example, we found that methane (CH_4) was oxidized readily in air to produce carbon dioxide and water.

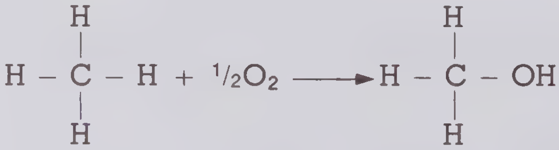
$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

This equation tells us that each methane molecule must gain two oxygen molecules in order to be completely oxidized to CO_2 and

H_2O. If, however, the methane gained only one oxygen atom the reaction may be something like this:

$$CH_4 + {}^1/{}_2O_2 \longrightarrow CH_4O$$

or graphically:



$$H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-H + {}^1/{}_2O_2 \longrightarrow H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-OH$$

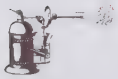
The addition of this oxygen atom to the methane molecule in this manner leads to a new series of organic compounds called *alcohols*.

Alcohols are characterized by the *OH* group as shown in the graphic formula. More generally then, substituting an OH group in a hydrocarbon molecule produces an alcohol.

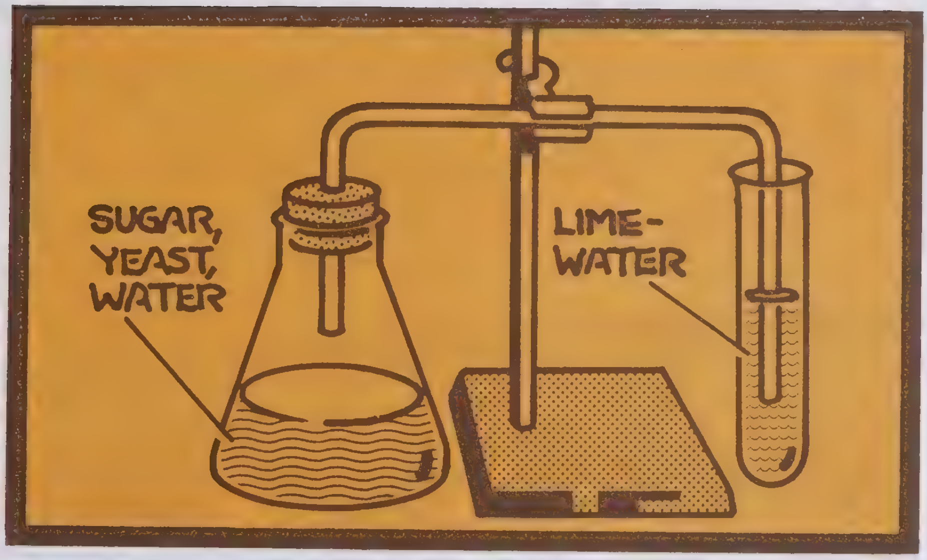
The alcohol produced from methane is called methanol. The alcohol produced by oxidizing ethane is ethanol. What name would be given to the alcohol derived from the next three alkanes after ethane? The -ol ending in future should indicate to us the presence of the OH group and hence an alcohol.

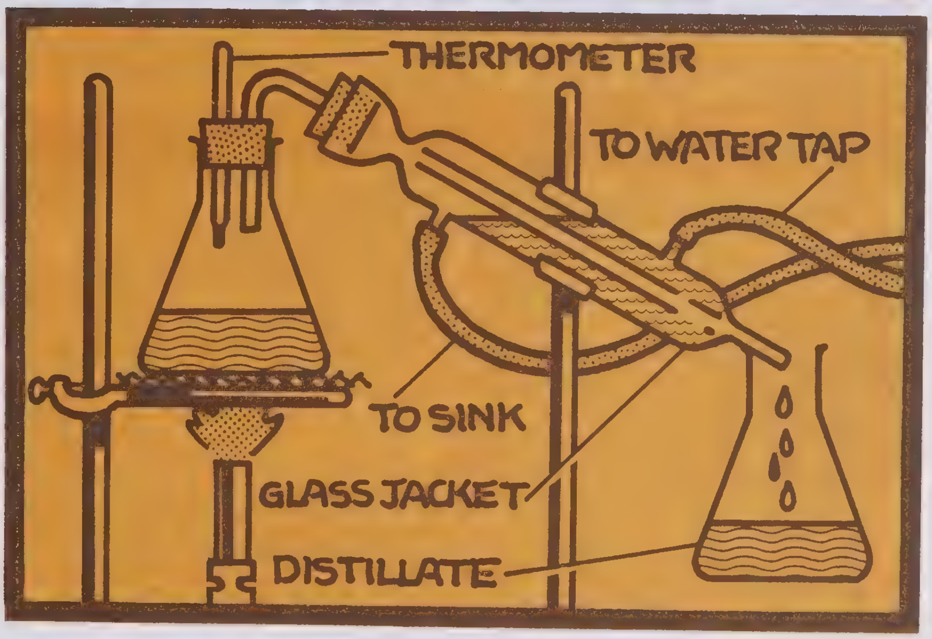
The OH group of an alcohol is called its *functional group* since chemical reactions of alcohols take place at this point in the molecule. The reactivity or the lack of reactivity of the functional group, then, determines much of the chemical behaviour of the molecule.

6.8 Making Ethanol by Fermentation



10 percent solution of glucose or corn syrup in water
yeast
limewater
250 ml erlenmeyer flask
one-holed stopper
delivery tube
test tube
retort stand and clamp
crucible
splint

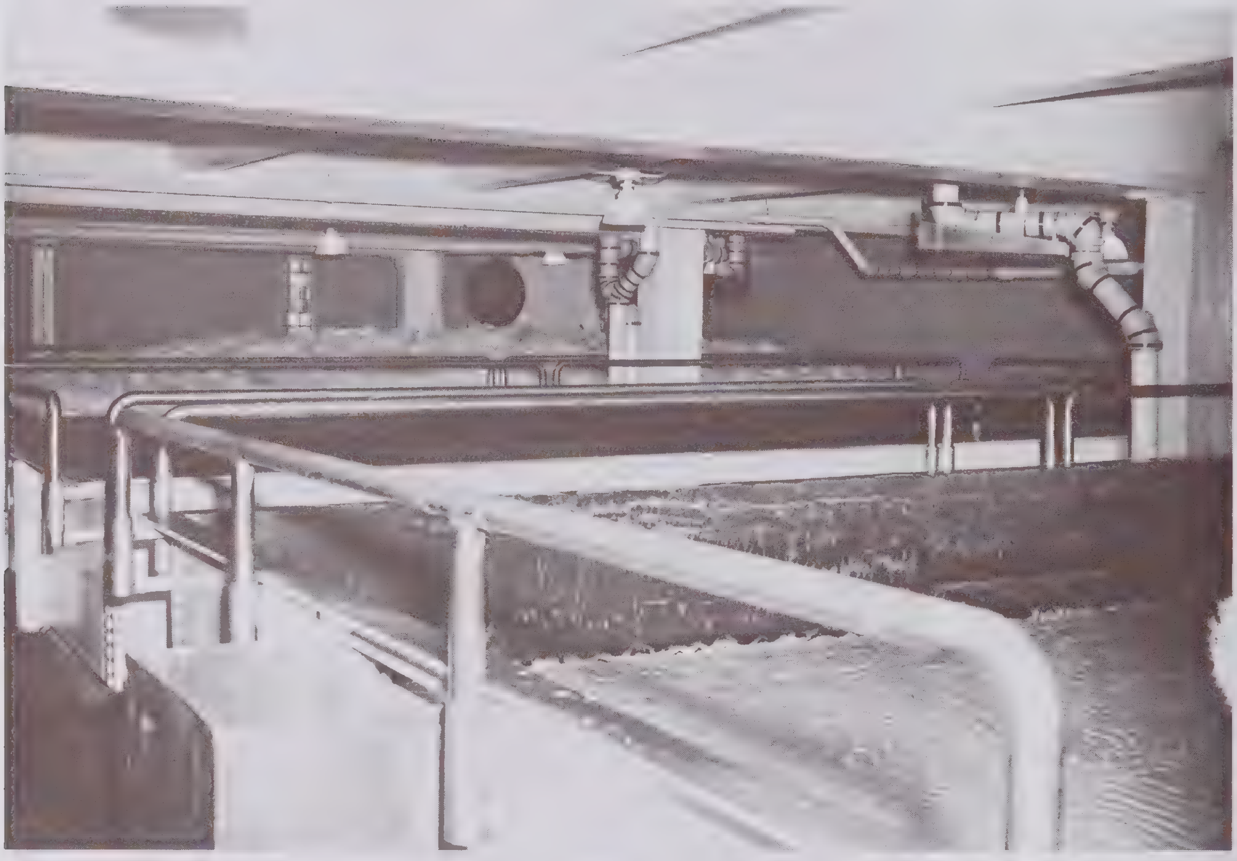

SUGAR,
YEAST,
WATER
LIME-
WATER


THERMOMETER
TO WATER TAP
TO SINK
GLASS JACKET
DISTILLATE

1. Make up the glucose solution or obtain about 50 ml of an already prepared solution and put it into a 250 ml erlenmeyer flask. Add a spatula measure of yeast and stir it into the solution.
2. Fit the flask with a one-holed rubber stopper and delivery tube. Place the delivery tube in a test tube half filled with limewater. Put the entire assembly in a warm area in the room for about two days.
3. After two days note the appearance of the limewater. Remove the stopper from the flask and filter about 20 ml of the solution from the flask. Describe the odour of the filtrate.
4. Put about 0.5 ml of the filtrate in a crucible and attempt to light it with a burning splint.
5. Pass the remaining filtrate to the teacher where it will be distilled.

6. Your teacher will distill off a few millilitres of the filtrate. Smell the distillate. What happens when a burning splint is brought to a small amount of the distillate?

The distillate should contain ethyl alcohol.



Brewery vat. The conversion of sugar into alcohol by yeast is a form of fermentation. The fermentation vats shown here are used by brewers to make beer. *Courtesy of Molson's Brewery, Ltd.*

Questions

1. What change(s) had taken place in the flask?
2. What change had taken place in the limewater? Explain this change.
3. Why did the distillate burn, but not the filtrate?
4. What name is given to the process that took place in the flask?

6.9 Some Reactions of Alcohols

ethanol
0.01 M $KMnO_4$
6 M H_2SO_4
6 M NaOH
methanol
copper wire
formaldehyde
4 test tubes
bunsen burner
tongs

1. Put about 2 ml of distilled water into the bottom of three small test tubes. To the first of these three test tubes add 2 ml of 6 M H_2SO_4 and to the second add 2 ml of 6 M NaOH.
2. Add 2 ml of 0.01 M $KMnO_4$ solution to all three test tubes and finally about 2 ml of ethanol to each of the three tubes. Shake each gently and observe any change in each tube.

$KMnO_4$ is a strong oxidizing agent. The clue that we may look for in this reaction to see if the alcohol has in fact been oxidized is a colour change in the permanganate ion. The colour of the solution may change from green to almost colourless depending upon the degree of oxidation.

3. Put 1 to 2 ml of methanol into a small narrow test tube. Note the odour of the alcohol.
4. Prepare a coil of copper wire by winding a strip of wire around a stirring rod. Leave about 15 cm for a handle.

Before proceeding make sure that the coil will fit easily into the test tube with the alcohol.

5. Hold the wire with your crucible tongs then heat it to red hot in your bunsen flame. Quickly lower the coil into the test tube and hold it just above the surface of the liquid. Describe the changes that take place on the copper. Cautiously note the odour of the vapours coming from the test tube.
6. Cautiously note the odour of the vapours from a bottle of formaldehyde.

Questions

1. What colour was the original permanganate solution?
2. What colour did the permanganate become in each of the three test tubes?
3. What was different about each tube to account for any differences in colour?
4. What condition(s) seems to favour the oxidation of an alcohol by $KMnO_4$?
5. What evidence was there that a chemical reaction was taking place between the hot copper coil and the alcohol?
6. Was the odour from the methanol different after the coil had been placed in the test tube from before it was put in the tube?
7. Did the odour resemble that of formaldehyde?

Aldehydes and Ketones

The reaction between heated copper and methanol was really one between hot copper oxide and methanol. The reaction is as follows:

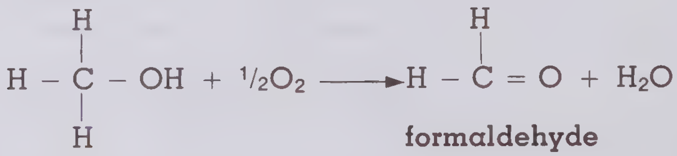
copper (II) oxide + methanol ⟶ copper + water + formaldehyde
$CuO + CH_3OH \longrightarrow Cu + H_2O + HCHO$

The evidence that we gathered during the experiment to indicate this reaction to us was the disappearance of the dark coating on the coil and the exposure of bright copper, as well as the distinct odour of formaldehyde, a liquid used to preserve biological specimens.

If we look more closely at the reaction, however, we will see that the alcohol has been oxidized. The alcohol itself is the product of the oxidation of a hydrocarbon. Organic compounds called aldehydes are the products of the oxidation of alcohols.

$CH_4 + {}^1\!/_2O_2 \longrightarrow CH_3OH$ Methanol produced by the oxidation of methane.

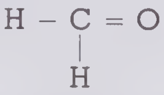
$CH_3OH + {}^1\!/_2O_2 \longrightarrow HCHO + H_2O$



$$\mathrm{H-\overset{\displaystyle H}{\underset{\displaystyle H}{\overset{|}{\underset{|}{C}}}}-OH} + {}^1\!/_2O_2 \longrightarrow \mathrm{H-\overset{\displaystyle H}{\overset{|}{C}}=O} + H_2O$$

formaldehyde

It is important to note here that the point of attack on the alcohol has been the OH functional group. The OH group characteristic of alcohols has been changed to a structure like



$$\begin{array}{ccccc} H & - & C & = & O \\ & & | & & \\ & & H & & \end{array}$$

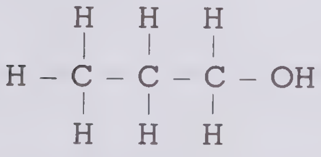
which is characteristic of a group of compounds called aldehydes.

The systematic naming of aldehydes is done in much the same manner as the naming of alcohols. The names are derived from the corresponding hydrocarbons having the same number of carbon atoms. The -ane ending is dropped from the alkane name and the ending -al is added. For example:
CH_4 is methane and HCHO is methanal;
C_2H_6 is ethane and CH_3CHO is ethanal;
C_3H_8 is propane and C_2H_5CHO is propanal.

It is interesting to note here, however, that a number of these organic compounds were known and named before systematic naming was developed. In some cases these names still remain and are more popular than the systematic name. Two examples of this arise in the aldehyde series with methanal which is better known as formaldehyde and ethanal which is better known as acetaldehyde. This could perhaps be compared to the person whose name is Arnold but for 75 years of his life has been called Max. It would be very difficult to drop the nickname and go to the full name of Arnold. The same sort of problem has arisen with naming of organic compounds.

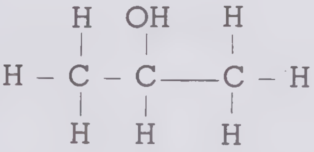
To add to the complexity of the naming and organizing of organic compounds, new compounds are formed by placing the functional group in different positions in the molecules. For example, in alcohols with three or more carbon atoms a new alcohol is formed by putting the OH functional group on the second carbon atom rather than on the first.

1-propanol:



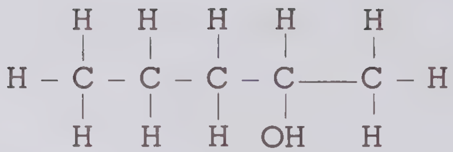
$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & OH \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

2-propanol:



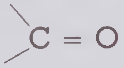
$$\begin{array}{ccccccccc} & & H & & OH & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & — & C & - & H \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

Both of these molecules are alcohols and, further, both are propanols. But due to the position of the functional group they have some different properties. Notice that the molecular formula is the same for each, C_3H_8O, but that the structural formula is different. Compounds with the same molecular formula but different structural formulae are called isomers. Numbers followed by a dash indicate the position of the functional group, as shown in the example above. The numbering is always done from the end of the molecule that will give the smallest number. For example, this is a molecule of pentanol.

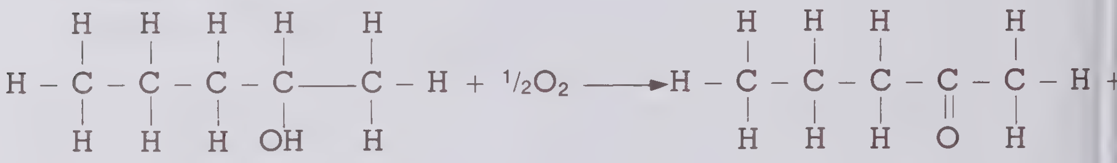


$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & OH & & H & & \end{array}$$

The five carbon atoms are indicated by the stem penta- (from pentane), and the OH functional group is indicated by the -ol ending. The position of the OH group must be indicated. Counting from the left puts the OH group on the fourth carbon atom. Counting from the right puts the OH group on carbon number two. Remember, always use the *smallest* numbering system. Therefore, the name of the alcohol represented above would be *2-pentanol*. The oxidation of an alcohol like 2-pentanol cannot yield an aldehyde since the carbon atom holding the functional group would not be left with a hydrogen atom which is necessary for the aldehyde functional group. A new series of compounds called ketones with the functional group



is formed when alcohols with the OH group on a carbon other than an end carbon are oxidized. This is the oxidation of 2-pentanol to produce a ketone.

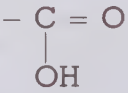


$$\begin{array}{ccccccccccccccccccccccccccccc} & & H & & H & & H & & H & & H & & & & & & & & H & & H & & H & & & & H & & \\ & & | & & | & & | & & | & & | & & & & & & & & | & & | & & | & & & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & H & + & \frac{1}{2}O_2 & \longrightarrow & H & - & C & - & C & - & C & - & C & - & C & - & H \; + \\ & & | & & | & & | & & | & & | & & & & & & & & | & & | & & | & & \| & & | & & \\ & & H & & H & & H & & OH & & H & & & & & & & & H & & H & & H & & O & & H & & \end{array}$$

2-pentanol **2-pentanone**

Carboxylic Acids

It is possible to carry our oxidation series one step further and produce yet a new series of organic compounds with the functional group

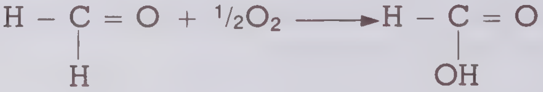


$$\begin{array}{cccc} - & C & = & O \\ & | & & \\ & OH & & \end{array}$$

The functional group is called *carboxyl* and the series of compounds with the carboxyl functional group is called *carboxylic acids*.

$HCHO + {}^{1}/{}_{2}O_2 \longrightarrow HCOOH$

formaldehyde ⟶ formic acid

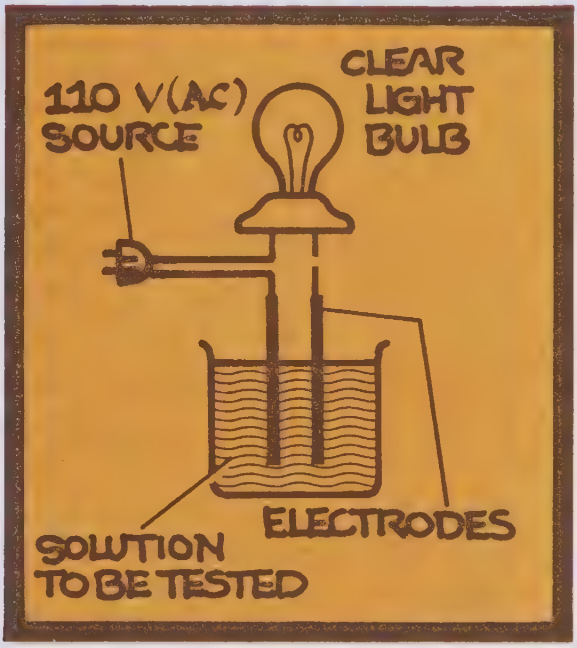


$$\begin{array}{ccccccccccccc} H & - & C & = & O & + & \frac{1}{2}O_2 & \longrightarrow & H & - & C & = & O \\ & & | & & & & & & & & | & & \\ & & H & & & & & & & & OH & & \end{array}$$

carboxyl group

6.10 Comparing Carboxylic Acids with Inorganic Acids

0.1 M acetic acid
0.1 M formic acid
0.1 M hydrochloric acid
0.1 M sulphuric acid
sodium carbonate
mossy zinc
red and blue litmus
bromthymol blue
light bulb
beaker
8 test tubes



1. To 1.5 ml of acetic acid solution in a test tube add a small amount of sodium carbonate (enough carbonate to cover the end of your spatula). Record your observations. Pay particular attention to the rate of the reaction.
2. To 1.5 ml of acetic acid solution in a test tube add a small lump of mossy zinc. Record observations.
3. Test a small sample of the acid solution with red and blue litmus. Add a few drops of bromthymol blue to a few millilitres of the acid solution.
4. Test the electrical conductivity of the acetic acid solution using the arrangement shown in the diagram.

Warning

Exercise extreme care with the bare electrodes.

5. Repeat Steps 1-4 with each of the other three acids.

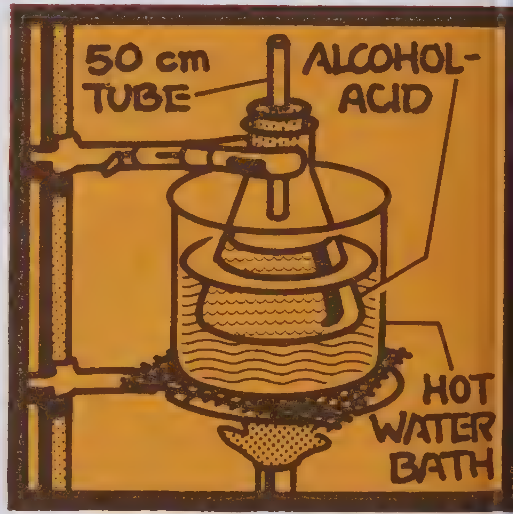
Questions

1. Describe what happens when sodium carbonate is added to an acid.
2. Describe what happens when zinc is added to an acid.

3. Does the rate of evolution of a gas from the carbonate-acid reaction or zinc-acid reaction have any dependence on the acid used? Explain.
4. Describe what happened to the litmus and to the bromthymol blue solution with each acid.
5. Write formulae for each of the acids used. What does each have in common that would account for any similarities in properties?
6. When acids are placed in water they tend to dissociate into ions which enables them to conduct an electric current. Some acids dissociate to a greater degree than others. What conclusion(s) would you draw about the degree of dissociation of the four acids used in this experiment?
7. What generalization could be made about organic acids compared to inorganic acids? (Formic and acetic acid are organic acids.)

6.11 Mixing Alcohols and Acids

ethanol
glacial acetic acid
concentrated sulphuric acid
beaker
50 cm glass tube
125 ml erlenmeyer flask
bunsen burner
retort stand and clamps
wire gauze
one-holed stopper



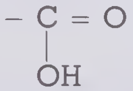
1. Select a 125 ml erlenmeyer flask for a reaction vessel and fit it with a one-holed stopper into which a 50 cm glass tube has been inserted.
2. Add 5 ml of ethanol and 6 ml of glacial acetic acid to the flask as well as 8 to 10 drops of concentrated (18 M) H_2SO_4.
3. Warm the mixture in the flask until it boils gently. The heating should be done in a hot water bath as illustrated in the diagram.
4. The long tube in the flask serves as a condenser. If droplets of liquid form in the tube and start to rise to the top, immediately remove the open flame.
5. Continue this gentle heating for about 15 min then remove the flask from the water bath and allow it to cool. Remove the condenser and examine the contents of the flask. Pay particular attention to the viscosity and the odour of the material remaining in the flask.

Questions

1. What changes have taken place in the mixture in the flask?
2. Describe the odour of the remaining contents of the flask. Can you name any common substance that has an odour similar to that of the final contents of the reaction flask?

Esters

Earlier we saw that the oxidation of an aldehyde resulted in the formation of an acid. All of these acids have the functional group COOH, called the carboxyl group. The atoms in the carboxyl group are bonded as shown below:



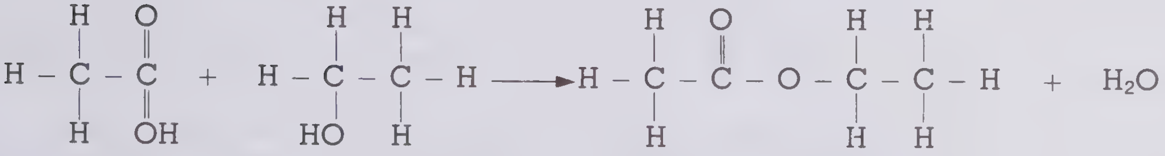
$$\begin{array}{l} -\,C = O \\ | \\ OH \end{array}$$

In aqueous solution the hydrogen may be separated from the carboxyl group as a hydrogen ion. This explains why organic molecules with the COOH functional groups have acid properties.

It is possible, however, to separate the OH group from the carboxyl group and replace it by other groups. The resulting molecules are called *acid derivatives*.

An acid derivative is characterized by a carbon atom that is not bonded directly to other carbon atoms or to hydrogen atoms.

One group of acid derivatives is formed by reacting organic acids with alcohols. The resulting derivatives are called *esters*. The reaction between ethanol and acetic acid is represented by the following equation:



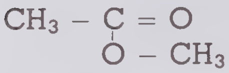
$$\begin{array}{ccccc} & H & & O \\ & | & & \| \\ H\,- & C & - & C \\ & | & & \| \\ & H & & OH \end{array} + \begin{array}{cccccc} & H & & H & \\ & | & & | & \\ H\,- & C & - & C & -\,H \\ & | & & | & \\ & HO & & H & \end{array} \longrightarrow \begin{array}{ccccccccc} & H & & O & & & H & & H \\ & | & & \| & & & | & & | \\ H\,- & C & - & C & - O - & & C & - & C\,-\,H \\ & | & & & & & | & & | \\ & H & & & & & H & & H \end{array} + H_2O$$

acetic acid **ethanol** **ethyl acetate** **water**

Notice that the carbon chain in ethyl acetate is separated into two parts by an oxygen atom. Ethyl acetate is an ester.

Esters are very common in nature. Esters are found in the essential oils of many plants and produce the odours in fruit. Esters of the low molecular mass acids and alcohols are used in perfumes and flavourings because of their particularily pleasant odours. Others are used as solvents for certain glues or nail polish.

Esters are named in accordance with the alcohol and the acid from which they are derived. The naming is done in much the same manner as salts are named. For example, the salt with the formula Na CH_3 COO is named sodium acetate. The name of the radical from the acid ends in -ate. In esters, the name of the acid part ends in -ate and the metal part of a salt is replaced by the name of the alcohol. The ester below is called methyl acetate.



$$CH_3 - \underset{\displaystyle O - CH_3}{\underset{|}{C}} = O$$

Select some other alcohols and acids and prepare a variety of esters. When you have prepared an ester allow other members of the class to see and examine it while you do the same with theirs. Try to associate the odour with some substance with which you already are familiar.

6.12 Preparation of an Ester

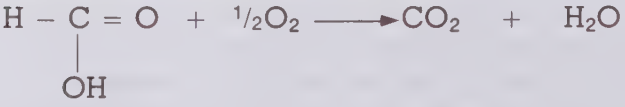
Alcohols and acids available to the teacher

1. Select one alcohol and one acid from those that your teacher has provided for you. Follow the same procedure as in Experiment 6.11. Remember to add the concentrated acid once again as it serves as a catalyst for the reaction. When you have prepared your ester allow other members of the class to examine it.

Questions

1. Write formulae for each of the alcohols used in this experiment.
2. Write formulae for each of the acids used in this experiment.
3. Combine the acids and alcohols in all possible combinations and write formulae and names for the esters that would result.

To complete our oxidation series we must now oxidize a carboxylic acid. Earlier we oxidized formaldehyde to produce formic acid so let is oxidize formic acid:



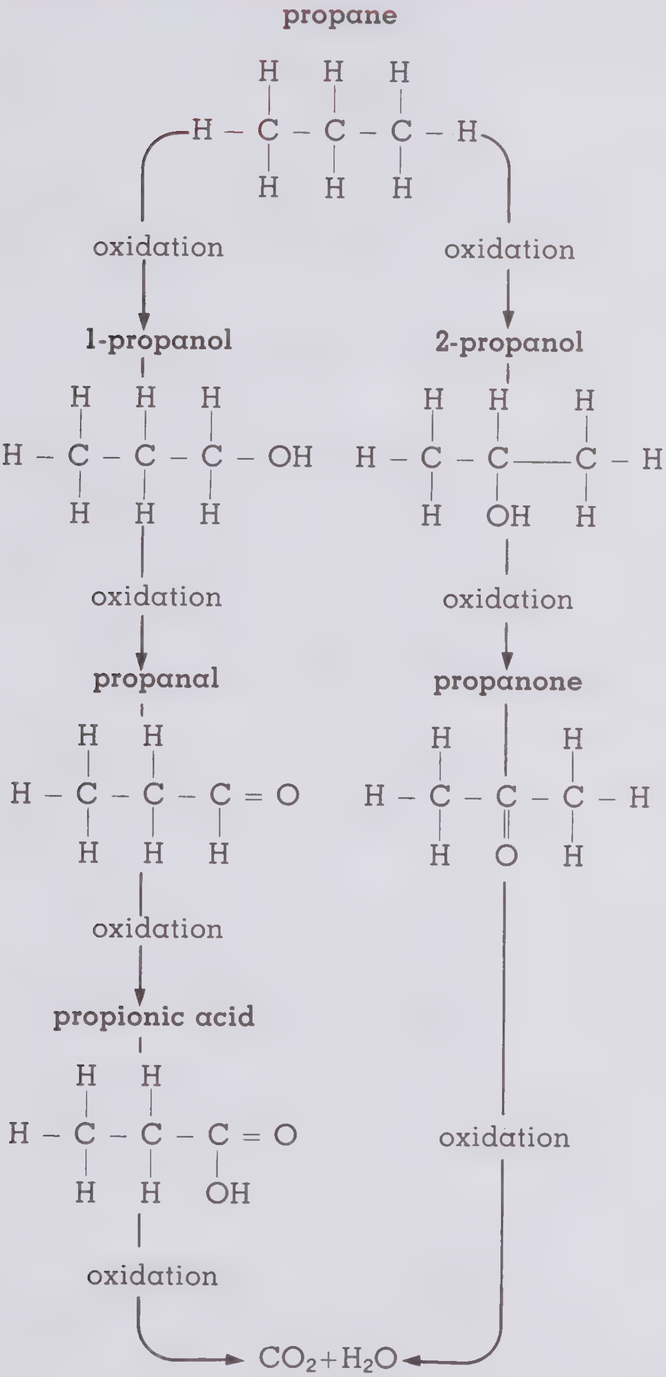
$$H - \underset{\displaystyle OH}{\underset{|}{C}} = O \;+\; {}^1/{}_2O_2 \longrightarrow CO_2 \;+\; H_2O$$

formic acid **carbon dioxide** **water**

The final stage in our oxidation is the formation of carbon dioxide and water.

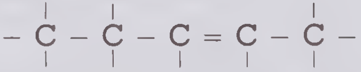
In summary then, the oxidation of a hydrocarbon produces an alcohol; the oxidation of an alcohol produces an aldehyde or ketone; the oxidation of an aldehyde produces a carboxylic acid and finally, the oxidation of a carboxylic acid produces CO_2 and H_2O. This oxidation series is illustrated in the table below by graphic formulae beginning with propane.

Oxidation series of a hydrocarbon to CO_2 and H_2O


propane
oxidation
oxidation
1-propanol
2-propanol
oxidation
oxidation
propanal
propanone
oxidation
propionic acid
oxidation
oxidation
CO2+H2O

Naming Organic Compounds

Close examination of the formulae in Table 6.2 reveals that within a series of compounds each molecule differs from the preceeding molecule by only one carbon atom. Just as carbon forms the basis of the molecule, it forms the basis for naming the molecule. Regardless of the series of organic compounds in question, the name is arrived at in much the same manner. First the number of carbon atoms in the chain is counted and the proper stem is used to identify that number of carbon atoms. A suitable ending is then selected to identify the series to which the molecule belongs. In many cases numbers are used to indicate the position of special groups or unique bonding arrangements.



$$-\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-\underset{|}{C}=\underset{|}{C}-\overset{|}{\underset{|}{C}}-$$

2-pentene

In naming the molecule above the number of carbon atoms is determined and the stem pent- is used to indicate five carbon atoms. The double bond makes the molecule an alkene thus the -ene ending. Finally, since there are different positions that the double bond could occupy, a number is used to indicate the position. Remember, numbering is always done from whichever end results in the smallest number.

The characteristic feature of this molecule is the carbon—carbon double bond. The characteristic reactions of the alkenes are those that take place at the double bond. The special bonding or group of atoms that defines the structure and determines the properties of a series of organic compounds is called the *functional group*.

Table 6.1

Series or Family	*Alkane*	*Alkene*	*Alkyne*	*Alcohol*
Functional Group	$-\overset{\vert}{\underset{\vert}{C}}-$	$-\overset{\vert}{C}=\overset{\vert}{C}-$	$-C\equiv C-$	$-\overset{\vert}{\underset{\vert}{C}}-OH$

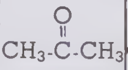
Series or Family	*Aldehyde*	*Ketone*	*Acid*
Functional Group	$-\underset{\underset{H}{\vert}}{C}=O$	$-\overset{\vert}{\underset{\vert}{C}}-\overset{\vert}{\underset{\underset{O}{\Vert}}{C}}-\overset{\vert}{C}-$	$-\underset{\underset{OH}{\vert}}{C}=O$

You must also understand that we have only grazed the surface in the naming of organic compounds. Carbon chains with as many as 70 carbon atoms are known. New molecules may result from chains becoming rings, branches from chains, branches from rings, positioning of functional groups, more than one functional group and the list goes on and on.

In our study of carbon and its compounds so far we have looked only at the lower molecular mass molecules. We want now to turn our attention to some of these long carbon chain molecules that were mentioned above — to those molecules with high molecular mass.

In our discussion of alkanes it was mentioned that the low molecular mass molecules are gases, the intermediate molecular mass molecules are liquids and after about 18 carbon atoms the compounds become solids.

Table 6.2
Regularities in Naming Organic Compounds

Number of Carbon Atoms	*Alkanes*	*Alkene*	*Alkyne*	*Alcohols*	*Aldehydes*	*Ketones*	*Acids*
1	CH_4 Methane			CH_3OH Methanol	HCHO Methanal		HCOOH Methanoic
2	C_2H_6 Ethane	C_2H_4 Ethene	C_2H_2 Ethyne	C_2H_5OH Ethanol	CH_3CHO Ethanal		CH_3COOH Ethanoic
3	C_3H_8 Propane	C_3H_6 Propene	C_3H_4 Propyne	C_3H_7OH Propanol	C_2H_5CHO Propanal	 $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}CH_3$ Propanone	C_2H_5COOH Propionic
4	C_4H_{10} Butane						
5	C_5H_{12} Pentane						
6	C_6H_{14} Hexane						
7	C_7H_{16} Heptane						
8	C_8H_{18} Octane						
9	C_9H_{20} Nonane						
10	$C_{10}H_{22}$ Decane						

Plastics

It should be apparent that if we can group together small molecular mass molecules to produce higher molecular mass molecules, then we should be able to produce solids. If these solids have certain desired properties such as high tensile strength or heat resistance they may be put to good use. This is exactly what has happened in organic chemistry and a group of substances called *plastics* has been developed. Plastics include a wide variety of useful materials with a multitude of uses. Plastics are used for things as simple as the casing on your pen, to the bodies of some automobiles.

These large molecules begin initially with molecules having only a few carbon atoms which are then bonded together time after time to produce a molecule with the desired properties. The small "units" that are repeated over and over again in one of these large molecules is called a *monomer*. The substance resulting from the union of the monomers is called a *polymer* and the process of producing a polymer is called *polymerization*.

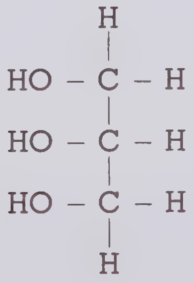
There are two ways by which small molecules can be linked with each other to form polymers. The first of these, called condensation polymerization, involves the elimination of water between the molecular pairs. The second is called addition polymerization in which small molecules with double bonds are involved. The double bond opens and the electrons form bonds with neighbouring molecules. In the experiments that follow we are going to prepare polymers by condensation polymerization.

6.13 Condensation Polymers

Part A

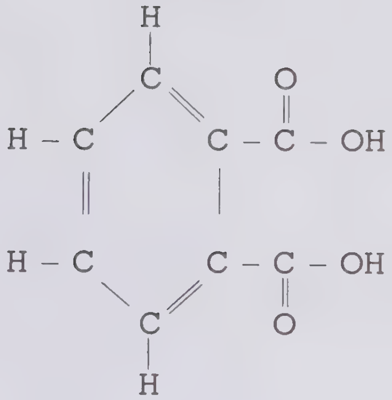
In order that a water molecule may be eliminated in the formation of the polymers, the monomers (the small units) that form the molecule must contain OH groups or at least O in order to contribute to the formation of water. For our first polymer we shall select glycerol and orthophthalic acid which are illustrated by the graphic formulae in the diagram.

glycerol

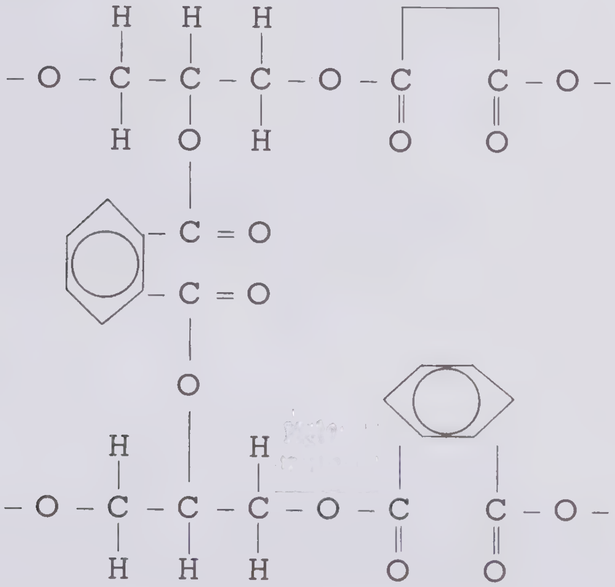


$$\begin{array}{c} \text{H} \\ | \\ \text{HO} - \text{C} - \text{H} \\ | \\ \text{HO} - \text{C} - \text{H} \\ | \\ \text{HO} - \text{C} - \text{H} \\ | \\ \text{H} \end{array}$$

orthophthalic acid



This partial formula shows how glycerol and orthophthalic acid could combine.



small baby food jar
glycerol
orthophthalic acid
watch glass
bunsen burner

1. Put about 2 g of glycerol and 3 g of orthophthalic acid in a small jar or beaker. Cover the beaker with a watch glass. Stir the mixture then heat over a *low* bunsen flame. Continue the heating until the bubbles puff up throughout the mixture. Allow the mixture to cool.

Warning

The contents of the beaker may burn.

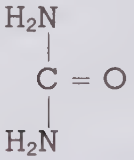
2. Remove some of the material from the jar and examine it then pass it on to your teacher. Your teacher will attempt to dissolve it in different solvents such as acetone, alcohol or water. Spread some of the acetone solution evenly on a piece of wood and then allow it to dry. Suggest a possible use for this polymer.

The polymer just produced is called glyptal resin. Paints made from resins similar to the one produced above are used for traffic markers; they are water resistant. The resin is also used in printing inks, enamels and furniture finishing. They are of great importance for coating purposes due to their permanence and good appearance.

Part B

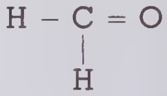
For our second condensation polymer we shall use urea, a constituent of urine which can now be made synthetically, and formaldehyde solution (formalin).

urea



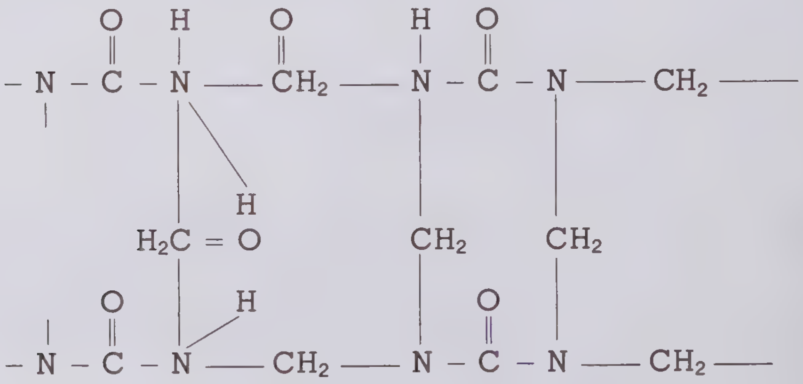
$$(H_2N)_2C=O$$

formaldehyde



$$H-\underset{H}{C}=O$$

This formula shows how the small molecules join to form the much larger molecule. The coloured portion shows how water molecules are formed in this reaction. All the atoms shown in colour have left the polymer to form water.



formaldehyde solution
urea
concentrated H_2SO_4
test tube
beaker

1. Put about 5 cm of formaldehyde solution into an 18 x 150 mm test tube and add urea with shaking until a saturated solution is prepared. Allow the excess urea to settle to the bottom then pour the clear solution into a small beaker.
2. Add about 7 drops of concentrated H_2SO_4 then stir the mixture. Describe the reaction that takes place. Is the reaction exothermic or endothermic?

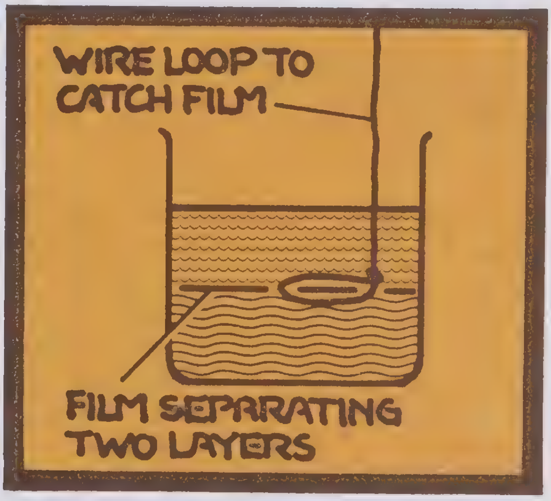
The material that you have just prepared is a plastic that may be shaped. After shaping, the plastic is baked whereupon it becomes a hard, brittle, heat-resistant substance that does not melt. Because it turns hard when it is heated and does not melt it is said to be a *thermosetting* plastic. It is used for radio cabinets, the casing for car batteries, the outer casing of telephones or electrical fittings and switches.

Other synthetic plastics may be reshaped by softening with heat. Plastic eyeglass frames are an excellent example of such a plastic. Plastics like this are called *thermoplastic*.

Part C

For our final condensation polymer we shall prepare nylon. This experiment should be demonstrated by your teacher.

carbon tetrachloride
sebacoyl chloride
sodium hydroxide
hexamethylene diamine
loop of copper wire
2 125 ml beakers
glass rod



1. Add 25 ml of carbon tetrachloride to a 125 ml beaker then dissolve into this 1 ml of sebacoyl chloride.
2. Into a second beaker prepare a solution consisting of 25 ml water and 1 g each of sodium hydroxide and hexamethylene diamine.

3. Carefully pour the water solution down the side of the first beaker so that it forms a second layer of liquid on top of the first. Look carefully into the beaker and you should see a thin film or skin between the two layers.
4. Make a loop on the end of a 15 cm length of copper wire. Put the loop into the mixture in the beaker down to the film and swirl the loop into the film; then slowly raise the wire from the beaker. Continue to do so until no further reaction occurs. Wind the fibre around a glass rod as it is drawn from the beaker.

Nylon is a thermoplastic since it can be melted and reshaped.

The molecules in the thread that was removed from the beaker are all jumbled and as a result the thread lacks tensile strength. The nylon is heated, then forced through small holes and drawn into thin threads. The molecules are all aligned by this process greatly increasing the tensile strength of the thread. A mass of individual pieces of string could be drawn into some kind of rope but if each string is braided with several others a much stronger rope is produced. The same thing happens when nylon is drawn into a thread.

Leo H. Baekeland

A plastic may be defined as a solid which may be moulded into a new shape. Such materials have been known and used for hundreds of years. Damp clay is a good example of such a material that has had many years of use. Other plastics that have been used for years come from trees or other plants. Shellac, for example, has been used for thousands of years in making moulded objects.

The plastics mentioned above occur naturally. Plastics as we know them, however, are synthetic or man-made.

The plastic industry really had its birth in the early part of this century. A Belgian born chemist named Leo H. Baekeland produced a synthetic plastic from phenol and formaldehyde.

Phenol is a crystalline substance that is soluble in water. Phenol is used extensively in many household disinfectants.

Formaldehyde is a gas with a very pungent, distinctive odour. Most often formaldehyde is used in solution form.

When phenol and formaldehyde were mixed nothing much appeared to happen. Baekeland boiled the mixture for some time and noted that the mixture separated into two layers — one colourless, the other yellow. The colourless layer was found to be water while the yellow layer was a gummy-like sap similar to that from a pine tree. Baekeland called this material a synthetic resin as opposed to a natural resin like shellac. The new resin appeared to be of little use initially but Baekeland found that when heated the yellow

Leo H. Baekeland

resin melted then gradually became putty-like and finally hard and rigid. When heated further the new substance did not melt.

This new substance that Baekeland had produced was named bakelite after its discover. Today it is used for such things as electrical insulators, radio cabinets or pipe stems.

The work of Baekeland stimulated much of the research into synthetic organic resins which led to the development of the many plastics that we know today.

Leo Baekeland was born and educated in Belgium but while still young, he moved to the United States. While in America he devoted his time to research which eventually led to his discovery of what is known today as bakelite. Prior to this discovery, however, he improved methods of developing photographic film. He produced a special emulsion of silver chloride that permitted much more rapid printing and allowed the substitution of artificial light rather than sunlight for developing.

Baekeland died at the age of 81 in Yonkers, New York.

Summary

The element carbon is by far the most unique and versatile element from the periodic table. Because of this, the study of carbon and its compounds has become a separate field of study called *organic chemistry*. This name has persisted through the years even though the name carbon chemistry describes the field more adequately.

Carbon, as an element, is found in more than one form in nature. Each form has a distinctive set of properties making each very useful for specific purposes.

Atoms of carbon have the ability to bond covalently with most of the common elements to form binary compounds. However, the most unique property of carbon by far is its ability to bond covalently with other carbon atoms to form straight chains, branched chains and even rings of carbon atoms. This unique property means there are many more compounds that contain carbon than compounds that do not contain carbon atoms.

The enormous amount of material to be learned in organic chemistry is made somewhat simpler for us by dividing each molecule into two parts:

1. The unreactive hydrocarbon part;
2. The much more reactive functional group.

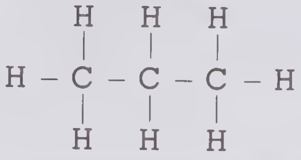
Molecules with the same functional group are placed in the same family of compounds. Each member of the family will have some different properties but many properties in common with all other members of the family. Rather than trying to study several thousand different molecules we can study about a dozen different families of compounds. By knowing some of the properties of a family of compounds we know some of the properties of each member of the family.

Finally, realizing that high molecular mass carbon compounds are solids and that low molecular mass compounds could be made to join under certain conditions has led to the discovery of a multitude of useful substances called plastics.

The process of joining the smaller molecules (monomers) into giant molecules is called polymerization. The field of polymerization has allowed man to design a plastic to suit almost every need.

Questions and Problems

1. What properties of the carbon atom make it unique and worthy of a separate field of study?
2. (a) Alkanes may be either gas, liquid or solid under normal conditions. Give the name and formula of an alkane that is a gas; a liquid; a solid.
   (b) How does the molecular mass of the solid alkane compare to the molecular mass of the liquid and the gaseous alkane?
3. Given is the graphic formula of propane (C_3H_8). Draw the graphic formula for:
   (a) an alkene;
   (b) an alcohol;
   (c) a ketone;
   (d) a carboxylic acid
   Give names for the four compounds you have drawn.



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

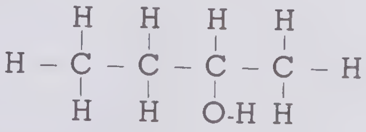
Remember that each carbon atom must have four covalent bonds.

4. Your text provides the names and formulae of only a few of the carboxylic acids. Draw graphic formulae and name the first five acids of this acid family.
5. Draw structural formulae for the functional groups of:
   (a) the alcohol family;
   (b) the ketone family;
   (c) the aldehyde family;
   (d) the ester family.
6. The position of a functional group is indicated by a number. The number indicates the carbon atom to which the functional group is joined.

Numbering is always done from whatever end provides the smallest number.

The compound below is called 2-butanol.

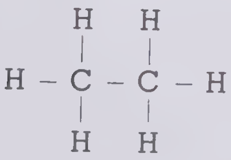
Account for:
(a) the number 2;
(b) the stem -but-;
(c) the ending -ol.



$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & O\text{-}H & & H & & \end{array}$$

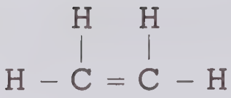
7. Name the following compounds:

(a)



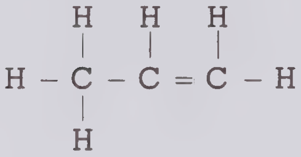
$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

(b)



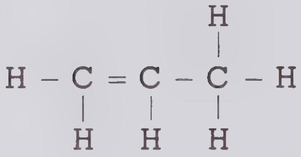
$$H-\overset{\overset{H}{|}}{C}=\overset{\overset{H}{|}}{C}-H$$

(c)



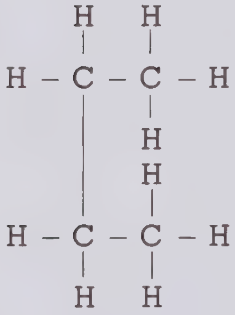
$$H-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-\overset{\overset{H}{|}}{C}=\overset{\overset{H}{|}}{C}-H$$

(d)



$$H-\underset{\underset{H}{|}}{C}=\underset{\underset{H}{|}}{C}-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-H$$

(e)



8. Draw structural formulae for the following compounds:
   (a) butane;
   (b) octane;
   (c) acetic acid;
   (d) propanal;
   (e) 1–pentanol.
9. Distinguish between a thermoplastic and a thermosetting plastic. Give one use for each.
10. Distinguish between a condensation polymer and an addition polymer.

Some Words You Should Know

distillation
destructive distillation
fractional distillation
alkane
alkene
alkyne
alcohol
aldehyde
ketone
functional group
carboxylic acid
ester
plastic
polymer
condensation
thermosetting
thermoplastic
oxidation

7

Water Hardness/
Soap and Detergents/
Bleaches/
Fragrances/
Distillation/
Photographic Film/
Properties of Metals/
Qualitative Analysis

Alka Seltzer
PHILLIPS

TUMS

Of Consumer Interest

To consume means to use up or destroy or waste away. All of us are destroyers of consumable goods every day when we eat our meals, wear our clothes, drive our cars, do our work or even in play. For years little thought was given to the limited supply of basic materials used to produce consumer goods. However, in the last few years, we have all become aware of the limited supplies of petroleum, the need for recycling waste materials and the necessity for reforestation. When you read any newspaper or magazine today you will probably find something about the need for wiser shopping or hints on how to save energy. We are continually reminded that there is not an unlimited supply of all the things we have taken for granted.

In this unit we want to show you how some consumables are made and how they can be identified. As this is a chemistry text, you will look at the chemistry involved in the tests and processes. Some of the fundamental concepts that were established earlier in the text will be extended to show you how chemistry can be applied to the production of consumable goods.

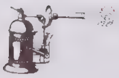
Hardness in Water

In many parts of the world there are extensive deposits of calcium carbonate on or near the surface of the earth. These may be limestone, chalk or marble. As rain falls on the deposits, rivers flow over them or underground water moves through them some of the calcium carbonate is dissolved to form a very dilute solution of calcium ions in the water. This experiment is intended to show you something of the properties of this natural solution.

This photograph of a heating coil in a water distillation flask shows some of the problems resulting from the use of hard water.

7.1

Properties of Hard Water



limewater
soap solution
sodium carbonate
drinking straw
4 test tubes

1. Since we do not have a solution of water known to contain calcium carbonate we will begin by making our own sample. Obtain a 40 ml sample of calcium hydroxide solution. This solution is saturated and is called limewater.
2. Pour 10 ml of this solution into a test tube. Label this Sample A.
3. Using a straw, blow your breath through the remaining 30 ml of solution until it turns cloudy. Divide this solution into three equal parts labelled B, C and D. Write an equation to represent the reaction that takes place between limewater and carbon dioxide. Explain why the solution becomes cloudy.
4. Using the straw, continue to blow your breath through Solution B until it becomes clear again. Write an equation to show the formation of calcium bicarbonate. Compare the solubility of calcium bicarbonate with calcium carbonate.
5. Very carefully warm Solution B to a gentle boil. Note the result. What would happen if water saturated in calcium bicarbonate were used in the family tea kettle or coffee pot? Write the equation for this reaction.
6. Add a single drop of soap solution to Solution A. Shake well. Note the result. Continue to add soap solution one drop at a time and continue shaking until you have a 5 cm deep layer of soap suds on the surface of the liquid. How many drops were needed to produce this lather?

7. Repeat Step 6 using Solution C.
8. Add about 0.5 ml of sodium carbonate to Tube D. Shake the test tube to dissolve the sodium carbonate. Repeat Step 6 using this solution.
9. Repeat Step 6 using a 10 ml sample of distilled water.

Questions

1. The curd or precipitate formed in Step 6 goes by many names. It is the very common "bath-tub ring" we see so often in TV ads. Chemically it is the same as a soap molecule except that the sodium in a soap molecule has been replaced by a calcium molecule from the solution. What is one difference between the sodium soap molecule and the calcium soap molecule?
2. Why is it so hard to form a soap lather in water containing calcium ions? What is the common name for this type of water?
3. How does sodium carbonate help to remove calcium from the water? Write an equation to represent the action of this washing soda on hard water.
4. Suggest several ways of determining if the water used in your home is hard water.

Supplementary Activities

With the approval of your teacher work out an experiment to answer some of the following questions.

1. Are any other substances besides calcium compounds found in natural hard water?
2. Do any substances besides calcium make water hard?
3. Do detergents behave in the same way as soaps in hard water?
4. Will any substance beside sodium carbonate soften hard water?
5. How does a home water softener work?
6. How is water softened in large quantities for industrial use?

In Unit 4 we discussed calcium and some of its properties. Calcium forms part of Group II on the periodic table along with beryllium, magnesium, strontium, barium and radium. This particular group of elements is often referred to as the *alkaline earth elements*. In our discussion of calcium, we found that it reacted fairly readily with water to produce hydrogen and a milky looking solution that appeared basic when tested with litmus paper.

$$Ca + H_2O \longrightarrow Ca(OH)_2 + H_2(g)$$

Calcium is one of the most abundant elements in the earth's crust, ranking fifth in abundance by mass. Since the element is so reactive it does not appear as the element itself but in various compounds, in particular calcium carbonate which appears in numerous forms such as limestone, marble or calcite.

When water soaking through the earth's crust reaches calcium carbonate or magnesium carbonate, chemical reactions occur to produce soluble compounds. The calcium ions or magnesium ions are dissolved in the water. It is very difficult to make a lather in this kind of water which is called hard water.

$CaCO_3 + 2H_2O + CO_2 \longrightarrow Ca\,(HCO_3)_2$

$Ca(HCO_3)_2$ is soluble.

The hardness in water is of two types, referred to as temporary or permanent hardness, depending on the ions dissolved in the water along with the metal ions. In temporary hard water there will be bicarbonate ions, HCO_3^-, and probably calcium, Ca^{++}, and magnesium, Mg^{++}, ions. When water with temporary hardness is boiled the bicarbonate ion breaks down into the carbonate ion, which in turn precipitates as calcium carbonate or magnesium carbonate.

$HCO_3^- + Ca^{++} \xrightarrow{\text{Heat}} CaCO_3(s) + H_2O + CO_2(g)$

If you live in an area with hard water you will probably have noticed a scale or deposit in the bottom of your tea kettle. This deposit is calcium or magnesium carbonate that has formed from the softening of the water each time it was boiled.

Permanent hardness in water, on the other hand, is not affected by boiling. But it is not a hardness that cannot be removed as its names seems to imply. Permanent hardness is the result of some ion, like the sulphate ion, in the water along with the Ca^{++} or Mg^{++} ions. Since the SO_4^{-} ion is unaffected by boiling, heating does not soften the water. Such water is softened by adding a soluble carbonate like sodium carbonate.

$Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{-}$

The dissolved Ca^{++} or Mg^{++} ions in the water will precipitate with the added CO_3^{-} ions.

$CA^{++} + CO_3- \longrightarrow CaCO_3(s)$

The Na^+ ion does not make the water hard as the Ca^{++} or Mg^{++} ions did.

Soaps, Detergents, Bleaches and Fragrances

7.2 Preparation of Soap

10 ml of oil (coconut oil)
15 ml of 5 M NaOH solution
salt water
filter
funnel
250 ml beaker
bunsen burner
long stirring rod
safety glasses

1. Put 10 ml of oil in a 250 ml beaker and then add 15 ml of 5 M NaOH.

Warning
NaOH solution is very caustic.

2. Gently heat the mixture with stirring. A *long* stirring rod should be used and safety glasses should be worn since spattering may occur when the reaction starts.
3. Continue heating until the substance in the beaker is a thick gelatinous mass. This may require from 10 to 15 min of gentle heating.
4. Remove the heat from the beaker and allow it to cool. Add 20 ml of cold salt water to the solid in the beaker and stir the mixture with your stirring rod. Decant the liquid and repeat this washing procedure twice more.
5. Remove the solid from the beaker with your stirring rod and press it into a filter cone that is supported in a funnel.

7.3 Properties of Soap

cottonseed oil
distilled water
soap from Experiment 7.2
3 test tubes
detergent
beaker
rubber band
dropper

1. Remove a small slice of the soap that you prepared in Experiment 6.14 and attempt to produce a lather by rubbing it briskly between your hands with some water.

Warning

Rinse your hands well after this procedure since your soap may still contain some NaOH.

2. Shake a slice of the solid soap that you have prepared with
   (a) 20 ml of tap water;
   (b) 20 ml of distilled water.
3. Float the rubber band on the surface of some water in a beaker. Carefully place one drop of soap solution inside the rubber band. Explain what has happened.
4. Half fill a test tube with water. Add 10 to 15 drops of oil. Note the relative positions of the oil and the water.
5. Shake the test tube vigorously. Again note the relationship between the oil and the water.
6. Add a few drops of soap to the test tube and shake again. Note the relationship between the oil and the water.

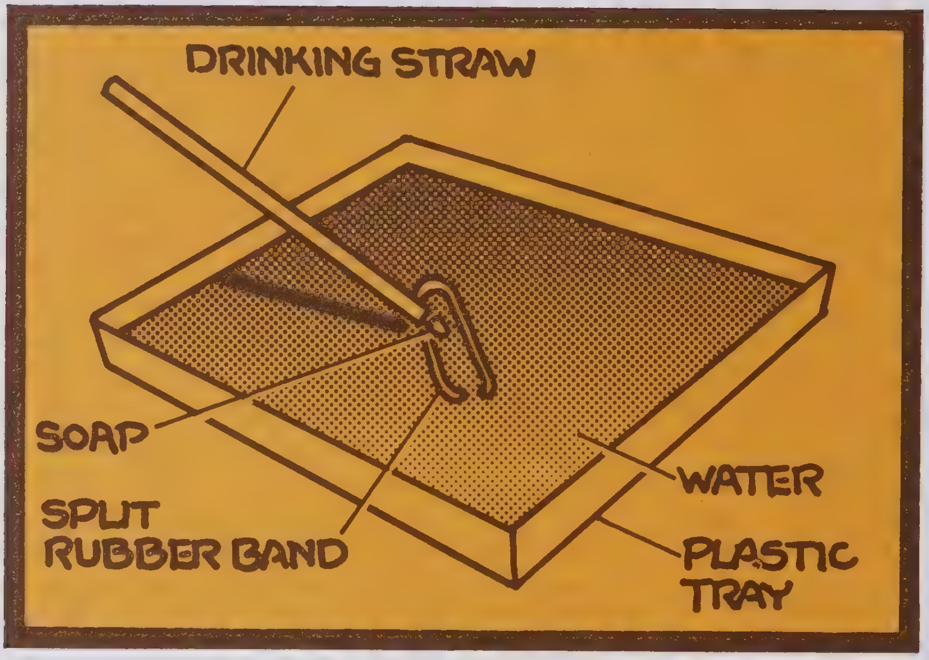
Questions

1. How would this action assist us in the washing process?

Fun and Games

Here is a chemistry joke to try on one of your friends. First rub a little soap on the end of your index finger without your friends seeing you. Now place a dish of water on the table so your friends can see it. Now for the story. You say, "Here is a swimming pool and here come the chemistry students." Now sprinkle the surface of the water with some household pepper out of a pepper shaker. Try to get the particles evenly distributed over the surface of the water. Now say, "Here comes the chemistry teacher." Now touch the tip of your soap covered finger to the centre of the surface of the water!

Well—it may not be all that funny, but there can be a *little* fun in chemistry!



7.4

Bleaching Action of Chlorine

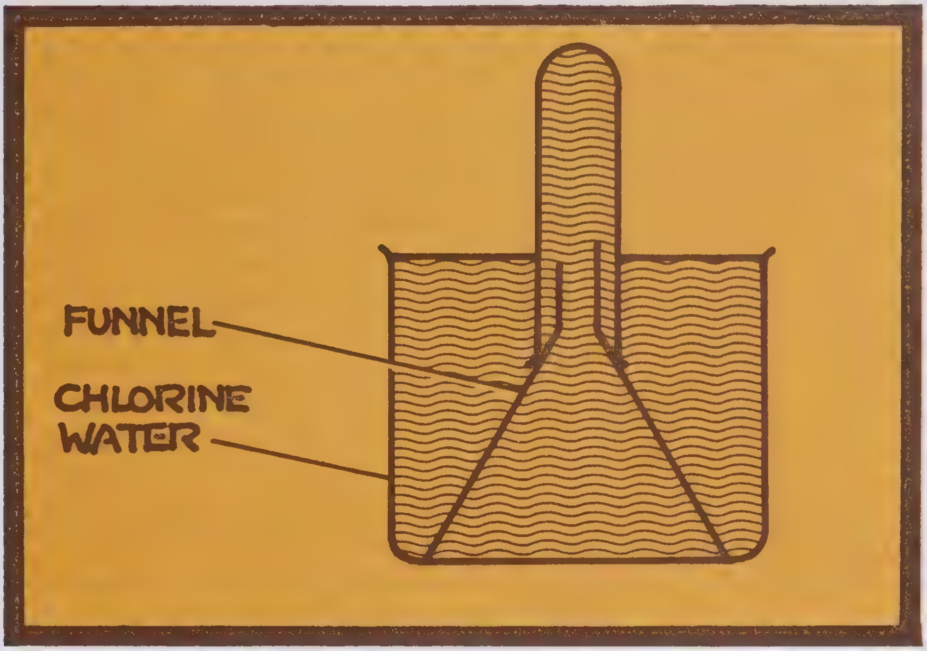
chlorine water
funnel
0.25 M NaOH solution
various samples of coloured cloth
beaker
test tube
splint

Part A Demonstration

1. Prepare about 500 ml of chlorine water by bubbling bottled chlorine through distilled water. Set up the apparatus as shown in the diagram and leave the assembly in the light for about a day. After a day note the change in colour of the chlorine water. Remove the test tube without losing the gas that has collected in it. Test this gas with a glowing splint.

Part B

1. Bring from home or obtain from your teacher a piece of coloured cotton cloth or even a piece of blue denim. First boil the cloth for a couple of minutes in a 0.25 M solution of sodium hydroxide. Using tongs, remove the cloth and rinse it in a beaker of water.
2. Dip the cloth into a beaker of chlorine water. When it is well moistened remove the cloth and dip it into a solution of 0.25 M sulphuric acid. Repeat the chlorine-acid treatment several times.
3. Rinse the cloth well under the tap to remove all excess chlorine and acid. Dry the cloth with paper towel.



Questions

1. Explain the disappearance of colour in the chlorine water in Step 1 of the experiment.
2. What gas collected at the bottom of the inverted test tube? How do you know? Where did this gas come from?
3. Sulphuric acid reacts with chlorine to liberate hypochlorous acid and sulphur dioxide. Write an equation for this reaction.
4. Dye stuffs react with an oxidizing agent to produce colourless compounds. What substance did the oxidizing in this experiment?

7.5 How Fragrances are Made

See Experiment 6.11, Mixing Alcohols and Acids.

The Manufacture of Soap

Soap as we know it can be classified into three main categories: household soap, toilet soap and industrial soap. In most cases, regardless of type, each contains additives to improve its cleansing ability and to improve its odour. In addition, practically all soaps contain anywhere from 10 to 25 percent water to improve their solubility.

The manufacturing process for soap is not very different from the process that you carried out in Experiment 7.2. The equipment is much more sophisticated but not the process. Once the fat mixture is melted, caustic soda solution (3.2 M NaOH) is added with stirring and heating. Small portions of the NaOH solution may later be added as long as the fat reacts with it. This is continued until the mixture becomes creamy smooth. At this point the soap undergoes what is called a graining process. This is done with about a 1 M NaCl solution. The undesirable products of the process called "lyes," which contain about 10 percent glycerol, are drawn off. After graining, the soap is heated once again with water and all last traces of fat are made to react with a very strong lye solution. After numerous washings with acid solution and water the soap is allowed to settle before being treated with dyes, perfumes or other additives. Soap made in this way is sold as bars, granules or flakes.

The chemistry involved in the manufacture of soap is not very different from what occurs when an acid reacts with a base.

Acetic Acid	Sodium Hydroxide		Sodium Acetate	Water
CH_3COOH +	$NaOH$	$\longrightarrow$	CH_3COONa	+ HOH

A soap is a salt of a special acid called a "fatty acid."

Stearin (Fat)	+ Sodium Hydroxide		Soap	+ Glycerol
$(C_{17}H_{35}COO)_3C_3H_5$ +	$3NaOH$	$\longrightarrow$	$3C_{17}H_{35}COONa$ +	$C_3H_5(OH)_3$

The fat in the equation above comes from the fatty portion of beef or lamb.

Sometimes substances called esters are added to the soap to give it a pleasant fragrance, especially to toilet soaps. Once again the chemistry of such a reaction is not very different from the acid-base reaction described above. When you made a fragrance in Experiment 6.11 you began with an alcohol and heated it with an acid.

Acid	**Alcohol**		**Reaction (Ester)**	Water
CH_3COOH +	C_2H_5OH	$\longrightarrow$	$CH_3COOC_2H_5$ +	H_2O

Substances like the one produced above give fruits their very own special flavour or the odour that one notices from a fruit like a pineapple. Fragrances prepared in this manner could be used as the perfume in the soap you prepared.

Bleaches

Any TV watcher knows from the advertisements that a soap alone will not remove stains like ground-in grass, blood or grease. A more powerful cleanser must be added to do this. Who likes to wear a new pair of blue jeans? Most people wash them several times before wearing them and often throw in a cup of bleach to help.

Bleaches are usually a solution of sodium hypochlorite ($NaClO$) or a solution of hypochlorous acid, ($HClO$). Both of these will decompose fairly easily, yielding atomic oxygen.

$$HClO \longrightarrow HCl + (O)$$

The atomic oxygen is very reactive and immediately reacts with dye, for example, to produce a colourless compound.

$$(O) + \text{dye} \longrightarrow \text{colourless compound}$$

In the bleaching experiment that you did earlier, the bleach used was a solution of chlorine and water.

$$Cl_2 + H_2O \longrightarrow HCl + \underset{\text{(hypochlorous acid)}}{HClO}$$

Note that the bleaching was not done by the chlorine specifically, but rather by the hypochlorous acid, which was a product of the reaction between chlorine and water.

Photography

7.6 Photographic Film

silver nitrate
sodium bromide
filter paper
dropper
2 test tubes
Bristol board 12 cm x 12 cm
paper towel

Part A Making the Film

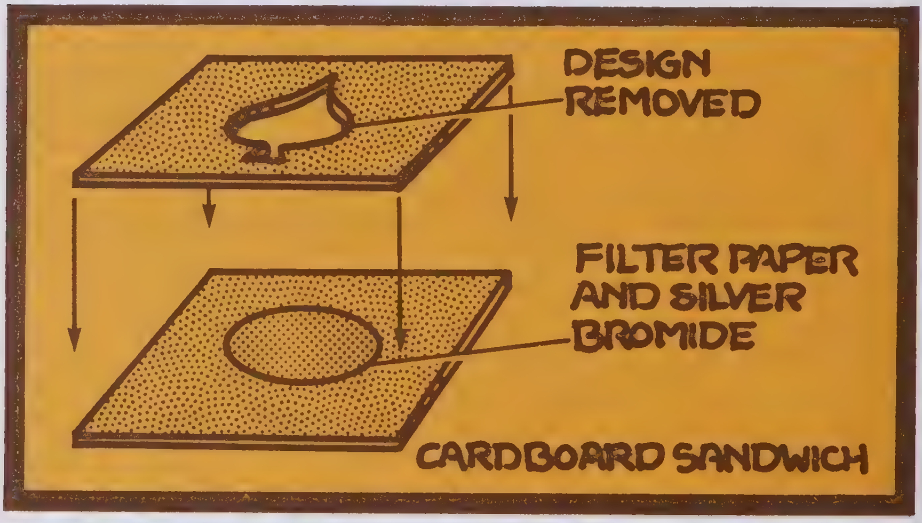
Work in a darkened room

1. Obtain a sample of sodium bromide in a small test tube. Place the filter paper on top of a folded sheet of paper towel. Using a dropper add enough sodium bromide solution to moisten the filter paper. (Do not soak the filter paper.)

2. Clean the dropper thoroughly or get a new one. Obtain a sample of silver nitrate in a second small test tube. Using a clean dropper, add the silver nitrate to the filter paper that has already been moistened with sodium bromide.
3. Press the filter paper between two dry sheets of paper towel to remove all excess water.

Part B Making the Negative

1. Obtain two 12 cm x 12 cm pieces of Bristol board or construction paper. Draw and cut out a design or picture on one board.


DESIGN REMOVED
FILTER PAPER AND SILVER BROMIDE
CARDBOARD SANDWICH

2. Make a sandwich of the two pieces of Bristol board and the treated filter paper. Make certain that the treated side of the filter paper faces the cut-out in the Bristol board.
3. Hold the sandwich for a moment toward a bright source of light. Going outside to the sunlight is best. The cut-out portion must be exposed to the light. In a dimly lit room separate the sandwich and note and record what has happened to the filter paper.
4. Allow the entire filter paper to be exposed to the light for a few moments and observe what eventually happens.

Questions

1. Silver nitrate and sodium bromide react with each other to produce silver bromide. Write a balanced chemical equation for the reaction.
2. Attempt to explain what happened to the exposed portion of the film you prepared.
3. Why must a camera film not be exposed to light before it is used?
4. Why is it necessary to rewrap a camera film after being used and return it to the canister for shipping?
5. Use a resource text to find out what is involved in:
   (a) developing;
   (b) fixing;
   (c) washing the film.
6. How is a print made from a developed film?

A film is made of a plastic-like material called cellulose acetate. This material is then coated with a light-sensitive material in a darkened room. This of course must be done because the film is not to be exposed to light until ready for use. The light-sensitive material is a silver compound much like the one you used in the experiment.

Once the film is prepared it can then be used in a camera. The film is first exposed to light when the photographer clicks the camera in taking a picture. The exposed film is then removed from the camera, again in a darkened room, and is developed. This involves soaking the film in certain chemicals to reduce any unchanged silver salts on the film. The film is then placed in a solution of hypo to dissolve any unchanged silver salt, leaving only metallic silver where light struck the film. This is called fixing. A piece of paper, treated like the film with a light-sensitive material, is placed beneath the film and exposed to light for a brief instant. Then it too must be developed, fixed and prepared for viewing.

Properties of Metals

How many metals do you know by name? How many uses for metals can you think of? If you sit down and consider these questions carefully you will probably find that you can name a dozen metals and give a hundred uses. Metals are the backbone of our technology. Metals form the skeleton of so many things that our world would stop if they were suddenly removed.

All the tall buildings and some not so tall would collapse in a heap. Every modern form of transportation would disappear. Every factory would be forced to shut its doors. Every power generating station, whether hydro, thermal or atomic powered, would cease to produce electricity. And we could go on and on and on.

Metals are important to our concept of civilization. We might ask why. Rather than considering the jobs they do, we will consider the properties for which they are selected for a specific job.

List a dozen uses of metals and beside each use, the property or properties that underlie that use.

Let us begin our study of metals by looking at some of these properties more closely.

7.7

Gardener's Sclerometer

30 cm piece of glass tubing
ruler
nail
several metal samples

1. Place a metal sample on a firm support and rest the open end of the vertically held glass tube on a flat surface of the sample.
2. Drop the nail down the tube and measure how high it bounces. Repeat the drop four more times in the same place. Average the bounce heights and then move the tube to a different surface or a different place on the same surface and repeat the drops.
3. Repeat the procedure on as many different metals as you can in the course of the period.
4. Use the nail to try and scratch each of the metals tested.

Questions

1. Why does the nail bounce higher on some metals than on others?
2. What property do you think is causing the nail to bounce?
3. Could you use this device to rate metals for this property?

4. Could you suggest a second property we might get information on from this apparatus if we modified it a little?
5. Explain why a knowledge of properties is important in the selection of a metal for a specific use.

Probably you decided that the sclerometer can give information about the hardness of a metal. There are many devices designed to do a better job than the sclerometer. (Few of them can compare for cost though.) The Birnell hardness tester is a common example. This device works by pressing a small object against the surface to be tested and then measuring the size of the dent made in the surface.

If we are only interested in which of two substances is harder we can apply the scratch test. If you run your fingernail across a candle then your nail scratches the candle. We infer that your nail is harder than the wax of the candle. There is a scale of hardness based on this idea of scratchability called the Moh Hardness Scale. Diamond, the hardest substance known, will scratch every other substance and is assigned a value of 10 on the scale. Other substances are assigned lower values.

10 diamond
9 corundum
8 topaz
7 quartz
6 feldspar
5 apatite
4 fluorspar
3 calcite
2 gypsum
1 talc

The important thing to note is that a substance higher in the scale will scratch all substances lower on the scale. It is therefore a hardness comparison scale. Nothing is actually measured.

Some metals are chosen for a job because they are hard. The metal used to make the jaws of a rock crusher or the metal used in the bearings in a wheel must be hard.

7.8 Tensile Strength (Demonstration)

This experiment will have to be a demonstration. You must record the results of it in the form of a graph.

0.05 cm copper wire
measuring tape
500 g masses
microscope
masking tape

1. Firmly attach a thin 0.05 cm copper wire to a very sturdy support. The lower end of the wire should have a loop in it.
2. Mark a section of the wire exactly 1 m long by wrapping it with pieces of masking tape. Carefully measure the distance between the tapes so it is exactly 1 m.
3. Add a 500 g mass to the wire and measure the length of the measured section. Record that length, subtract the original length and record the extension in mm/m. Continue to add 500 g masses and record the data in a table like Table 7.1

Table 7.1

Masses (g)	*Total Length (mm)*	*Extension (mm/m)*
0	1000	0
500		

4. Continue adding masses until the wire breaks. Examine the break under a microscope.
5. Plot a graph of masses versus extension.
6. Determine the diameter of the wire.

Questions

1. Why did the wire break?
2. How much mass could the wire have supported without breaking?
3. Calculate the mass that a wire of 1 cm^2 cross-section could have supported. The number of grams (or kilograms) that 1 cm^2 of wire can support is called the *tensile strength*. Table 7.2 gives the tensile strengths of some other materials.

Table 7.2

Materials	*Tensile Strength* (kg/cm^2)
chromium steel	5 500
manganese steel	8 000
stainless steel	7 400
aluminum (pure)	1 000
aluminum bronze	6 200
molybdenum (pure)	32 000
tungsten (pure	49 000
antimony lead	300
borosilicate glass	800
wood (with grain)	80
wood (across grain)	25

As you can see, there is a great variety in the tensile strength of materials. Many materials are strong in just the opposite way.

Instead of being able to resist being pulled apart, they are able to resist being compressed. Concrete is very strong in compression but rather weak in tension. Concrete is often prestressed by putting steel rods through the mixture. The steel then takes the tensile loads while the concrete takes the compressive loads. This combination of steel and concrete has proved to be very versatile and is widely used in the building industry.

Qualitative Analysis

The branch of chemistry referred to as analytical chemistry is divided into two major branches. One of these is *quantitative* analysis, while the other is *qualitative* analysis. In quantitative analysis the relative or often the absolute amount of one of the constituents of a material sample is determined. For example, an ore of nickel may be analyzed quantitatively to determine the percentage of nickel in any given sample. Perhaps on the basis of this analysis one might determine whether or not it would be economical to mine the ore. Qualitative analysis must precede a quantitative analysis since it only identifies a substance as being present. The results of a qualitative analysis are usually expressed in words or formulae of compounds or ions, but the results of a quantitative analysis are expressed as a number. In the nickel ore example given above the qualitative analysis would identify nickel as being present whereas the quantitative analysis would indicate how much nickel was present.

In the work that follows you will be first introduced to the concept of qualitative analysis and you will learn how to identify certain elements or ions. When the concept and certain tests are understood, you will carry out a qualitative analysis on a fertilizer and a laxative in an attempt to prove or disprove that there are certain elements in them.

7.9 Flame Tests

platinum or nichrome wire
bunsen burner
concentrated hydrochloric acid
solutions of various metal salts

1. Clean a platinum or nichrome wire by heating it in a hot bunsen burner flame and then plunging the hot wire into concentrated

hydrochloric acid. Repeat this until the hot wire does not colour the flame around it.

2. Dip the clean wire into a solution of copper sulphate. Hold it in the flame and note the colour. Repeat the dipping if necessary to establish the colour.
3. Clean the wire again and repeat the procedure using solutions containing lithium, potassium, barium, calcium, strontium and sodium. Do sodium last.

Questions

1. How could this test be used to identify certain metals?
2. Could you use this test to identify cobalt? Why?

7.10 Qualitative Analysis 1

6 M HCl
0.1 M K_2CrO_4
6 M NH_4OH
6 M HNO_3
lead nitrate solution
silver nitrate solution
mercuric nitrate solution
3 test tubes
beaker

1. Into a clean test tube place 2 ml of lead nitrate solution.
2. Into a second test tube place 2 ml of silver nitrate solution and into a third test tube place 2 ml of mercuric nitrate solution.
3. Add 20 drops of 6 M HCl to each test tube. Record the results.
4. Place the three test tubes in a beaker of boiling water for a few moments. Record the results.
5. Take 10 drops from the test tube that had the greatest solubility and add it to 1 ml of 0.1 M K_2CrO_4.
6. Take the three test tubes out of the hot water and cool them back down to room temperature.
7. Add 2 to 3 ml of 6 M NH_4OH solution to each test tube.
8. Add 3 to 4 ml of 6 M HNO_3 only to those test tubes in which the precipitate dissolved.
9. Dispose of the contents of each test tube as your teacher directs. Wash your hands carefully.

Questions

1. How could you use this test to identify silver? mercury? lead?
2. Write the equation for the initial reaction with the HCl.
3. Ask your teacher to give you a solution containing two of these substances. See if you can identify them from this procedure for silver, mercury and lead.

7.11 Qualitative Analysis 2

H_2S solution
solutions of various metal salts such as:
zinc
nickel
tin (II)
arsenic
cadmium
antimony
copper (II)
iron (II)
6 test tubes

1. Set up six clean test tubes with 2 ml of one of the following solutions in each:
   (a) zinc chloride
   (b) copper (II) sulphate
   (c) antimony (III) chloride
   (d) cadmium (II) nitrate
   (e) arsenic (III) chloride
   (f) tin (II) chloride
   (g) any other solution your teacher may direct you to use.
2. To each test tube in turn slowly add H_2S solution a drop at a time, carefully noting the results. Stop when you feel nothing further will be accomplished.

Warning

H_2S vapours are dangerous. Do not breathe the vapours.

Questions

1. How could you use this test to identify aluminum? chromium? zinc? manganese? iron? cobalt? nickel?
2. Write equations for the reactions with H_2S.

If you look back over the previous experiments you must realize that we have been developing a series of tests to apply to an unknown solution in order to try and identify it. At the moment we are capable of identifying the following substances in solution:

zinc
nickel
iron (III)
copper
antimony
cadmium
arsenic
tin
silver
mercury
lead
lithium
potassium
barium
calcium
strontium
sodium

There are 17 substances in this list. There are many more substances we do not have tests for but these should be enough to give us an idea of how we might analyze a substance.

The basic problem is two-fold:

1. Minerals are solids as a rule not solutions.
2. Minerals usually contain more than one metallic ion.

The first problem is fairly easy to overcome. We must dissolve the mineral in some substance such as nitric or hydrochloric acid. Sometimes it is necessary to use other solvents but these two are fairly common.

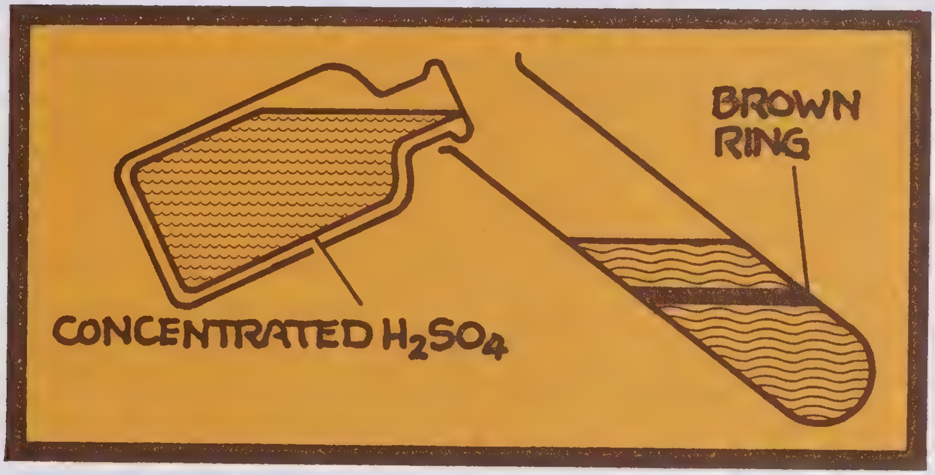
The second part of the problem is more complicated. It requires us to develop a scheme for analysis. Basically the scheme starts just as our experiments did.

The unknown solution containing the ions to be identified is first treated with hydrochloric acid, as in Qualitative Analysis 1. If lead, silver or mercury are present they will precipitate. The solution is filtered and the solid residue is treated to determine which of the three it is, as in Qualitative Analysis 1. The filtrate is now treated with H_2S solution as in Qualitative Analysis 2. Certain other substances may appear here and be identified. The solution is filtered and the filtrate is treated with NH_4OH as in Qualitative Analysis 1. Certain other substances may appear and be identified. Flame and borax bead tests may be used at any point to confirm or identify other substances.

This procedure is really a summary of the kind of techniques that must be used to identify the metals found in a mineral. Remember too, that our scheme does not include all the elements but we only want you to get the basic idea of how these things are done. You are not expected to be able to repeat this procedure by memory, nor will you be asked to state the procedures on a written examination. The examination on analysis is a practical one.

7.12 Qualitative Analysis 3

0.1 M KNO_3 solution
0.1 M Na_3PO_4 solution
ferrous sulphate solution (freshly prepared)
concentrated sulphuric acid
ammonium molybdate solution
test tube
beaker



Part A Nitrates

1. To 10 ml of the nitrate solution in a clean test tube, add 5 ml of the freshly prepared solution of ferrous sulphate. While holding the test tube at about a 45° angle, slowly add 5 ml of concentrated sulphuric acid. A brown ring between the layers in the test tube results from the presence of a nitrate. Since nitrates contain nitrogen, this test could be used in qualitative analysis to identify nitrogen.

Part B Phosphates

1. Boil about 15 ml of the phosphate solution for about 10 min or until the volume has been reduced to about one-third of the original volume. Filter the solution that remains, collecting only 2 to 3 ml of the filtrate. To the filtrate add about 10 ml of the ammonium molybdate solution. Warm the filtrate and molybdate mixture slowly in a beaker of warm water. A yellow precipitate indicates the presence of the phosphate ion.

7.13

Testing Fertilizers for Elements Necessary for Plant Growth

fertilizer sample
hydrochloric acid
ferrous sulphate
concentrated sulphuric acid
nichrome wire
cobalt glass
100 ml beaker
bunsen burner
3 test tubes
3 beakers

Part A

1. Add 5 g of the fertilizer sample selected to 15 ml of 0.5 M HC1 solution in a 100 ml beaker and then heat the mixture to boiling point. Filter the liquid and catch a few millilitres of the filtrate in a small test tube.
2. Do a flame test on the filtrate from Step 1, viewing the flame through a cobalt glass.

Questions

1. What element(s) were you able to identify by the flame test?
2. Use a reference text on plant growth to see if the element(s) identified are necessary for plant growth.

Part B

1. Dissolve about 3 g of the fertilizer under examination into 20 ml of distilled water. Shake the tube to dissolve the fertilizer. After several minutes filter about 10 ml of the solution into a clean test tube.
2. Do the brown ring test on the filtrate.

Questions

1. Were you able to identify the presence of nitrogen in the fertilizer in the form of a nitrate?
2. Does nitrogen play a significant role in the development of a plant? (Use a reference book.)

Part C

1. Dissolve 3 g of fertilizer in 20 ml of distilled water. Bring the solution to a boil and continue to boil for about 10 min. You do not want to boil the tube to dryness. After boiling, filter the solution and catch the filtrate in a small clean test tube. Only a couple of millilitres of the filtrate will be required.
2. Do the test for phosphates on the filtrate sample.

Questions

1. Were you able to identify any phosphates in the fertilizer that you have been studying?
2. Are phosphates necessary for the development of a healthy plant? (Use a reference text.) What part of the plant are they necessary to?

7.14 Magnesium in Laxatives

samples of solid laxatives
6 M ammonium hydroxide
6 M ammonium chloride
6 M sodium hydrogen phosphate
filter paper
mortar and pestle
2 test tubes

1. Grind a laxative tablet in your mortar and pestle. Transfer the powder into a clean test tube containing 15 ml of distilled water. Cork the tube and shake vigorously to dissolve the powder.
2. Filter the mixture in the test tube, catching the filtrate in a second clean test tube.
3. Add to the filtrate a few drops of ammonium hydroxide. If a precipitate does not form, discard the mixture and try a different laxative. If a precipitate does form allow it to settle in the test tube. Decant the liquid but save the solid material.
4. Add ammonium chloride, drop by drop, to the precipitate until the solid material has dissolved.
5. The final step in the test is to add 5 ml of sodium hydrogen phosphate to the clear liquid. If magnesium is present a white precipitate should appear.

7.15 Identifying Fibres in the Textiles used in Clothes

Samples of cloth. These should be supplied by the students. Rather than each student working with cloth he has provided himself, all samples should be pooled to make the following experiment more of a challenge.

splint
acetone
8 M HCl
6 M H_2SO_4
5.25% sodium hypochlorite solution (laundry bleach)

1. Each student should select from the collection of materials several samples of cloth that appear to be different. If microscopes are available each material selected should be examined under them. Threads from each sample should be examined.
2. Hold single fibres from each cloth in turn and ignite them with a burning splint. Note the rate of burning and the amount of residue left after the burning. Compare the burning results with Table 7.3.
3. Test the solubility of thread of the materials that you have selected in the solutions suggested below:

Warning
Acetone vapour is very flammable.

   (a) acetone
   (b) 8 M HCl
   (c) 6M H_2SO_4
   (d) 5.25 percent sodium hypochlorite solution (laundry bleach).
4. Compare the solubility of each fibre with Table 7.3 in an attempt to identify the fibre. The fibres must be left in each solvent for at least 15 min.

Table 7.3

Fibre	*Solubility*	*Burning*
Silk	The fibre dissolves in (a) 6 M H_2SO_4 (b) 10 M HCl (c) household bleach The fibre should be left in each of the above solvents for about 15 min. Silk does not dissolve in household bleach.	Odour something like burning hair. The material will not draw back from the flame as it burns.
Acetate	The fibre dissolves in (a) 6 M H_2SO_4 (b) 10 M HCl (c) acetone The fibre should be left in each solvent for about 15 min. Acetates do not dissolve in household bleach.	Material draws back from the flame as it burns. A hard brittle material is left after burning. It can be crumbled fairly easily.
Nylon	The fibre dissolves in (a) 10 M HCl (b) 6 M H_2SO_4 It does not dissolve in household bleach or acetone.	Material draws back from the flame as it burns forming hard balls of material that are difficult to crush.
Cotton	The fibre dissolves quite readily in (a) 6 M H_2SO_4. The fibre should be left in the solvent for about 15 min. Cotton does not dissolve in bleach or 10 M HCl.	Material does not draw back from the flame nor does it form balls of material like acetate or nylon.
Rayon	The fibre dissolves in (a) 6 M H_2SO_4 (b) 10 M HCl.	Burns in an identical manner to cotton.

8

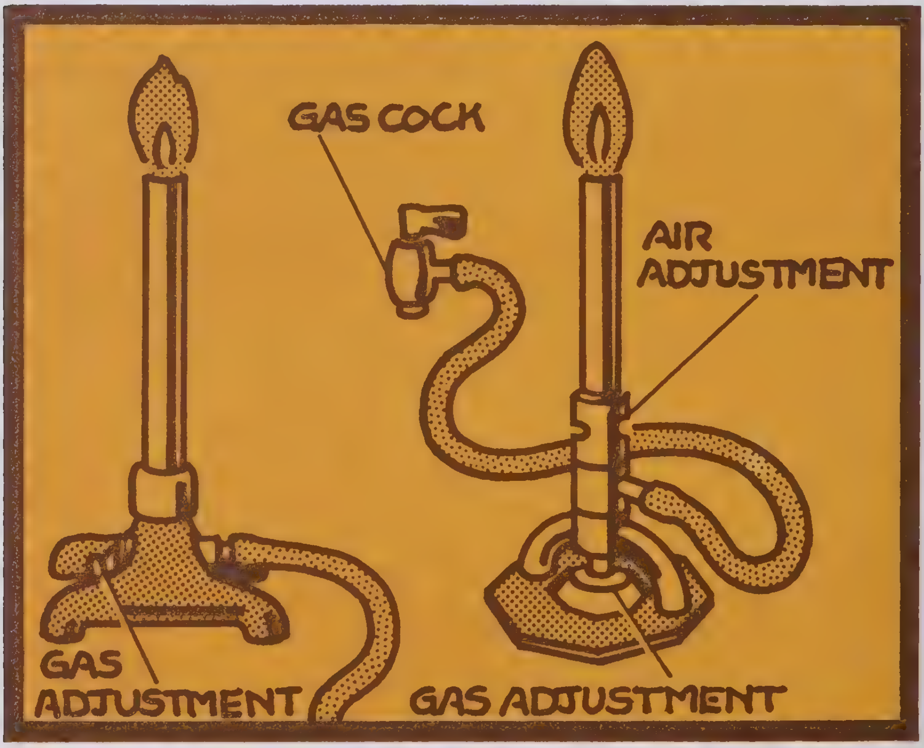
Bunsen Burner/Massing/Cleaning Glassware/Filtration/Measuring Volume/Glass Manipulations/Handling Liquids/Identifying Odours/Melting Point Determinations/Heating in a Crucible

Distilled
Water

Using the Bunsen Burner

Probably the best known piece of equipment in the laboratory is the bunsen burner. Simple as it is, it does require some knowledge if you are to get the best use out of it.

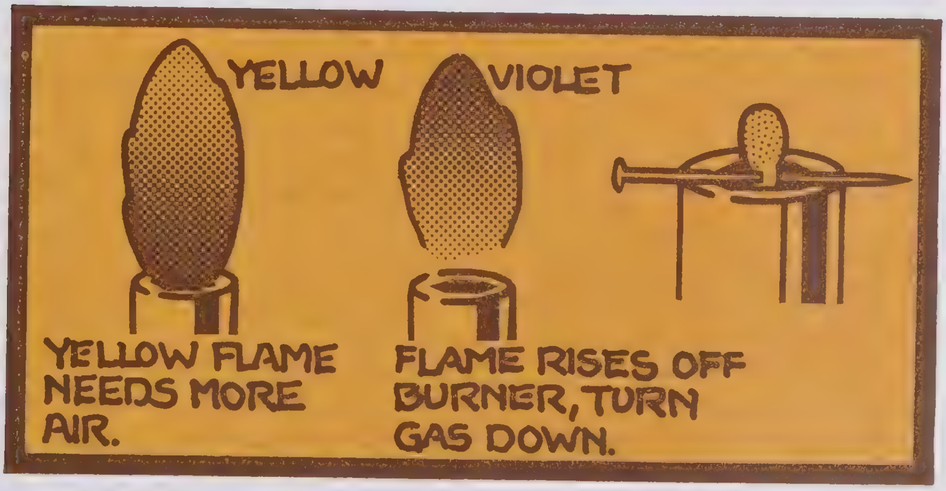
There are many designs of burners but the two most common are shown in the diagram on this page. In the burner on the right the gas flow is regulated with the knurled knob below the burner; in the other burner it is regulated with the bench tap. In the first the air supply is regulated by turning the tube or barrel of the burner up or down; in the second it is regulated with the lever beneath the burner. All burners have these two features — a gas regulating mechanism and an air regulating mechanism. Even if your burner is of a different type it will have these features.



8.1 The Bunsen Burner

bunsen burner
lighter
matches
pin
platinum or nichrome wire

1. Examine your burner. What type is it? Can you see how to regulate the air supply? The gas supply?
2. Connect the rubber hose to the burner and to the bench outlet. Make sure there are no kinks in the hose.
3. Close off the air supply regulator.
4. Light the match.
5. Turn on the gas and light it. (If your burner is of that type you must open both the bench tap and the burner tap.)
6. Slowly open the air control until the yellow smoky flame changes to pale blue. Draw a diagram of this flame.
   Here are some tips for adjusting the flame:
   (a) If the flame is yellow open the air valve.
   (b) If the flame seems to rise up off the barrel of the burner turn down the gas supply a little.
   (c) If the flame seems to roar turn down the air supply.
   (d) If there is a pop followed by a fluttery sound turn off the gas, wait a few moments for the barrel to cool down and relight starting at Step 3. (This problem is called *striking back* and is caused when the gas burns inside the barrel as well as above it.)
7. Practise adjusting the burner to give:
   (a) a small blue flame;
   (b) a large blue flame.


YELLOW
VIOLET
YELLOW FLAME NEEDS MORE AIR.
FLAME RISES OFF BURNER, TURN GAS DOWN.

8. Turn off the burner. Place a paper match in position as shown in the diagram. Relight the burner starting at Step 3.
9. Turn off the burner and remove the match. Relight starting at Step 3.
10. Hold the platinum or nichrome wire in the flame for several seconds. Note the results. Move the wire to a different region of the flame and repeat. Draw a temperature map of the burner flame.
11. When you have completed the experiment turn off the gas. Remember to turn off the gas at the bench if your burner has a gas adjustment.

Questions

1. Compare the diagram of the flame with the temperature contour map. Do you think colour is a guide to temperature?
2. Why does the match behave as it does?
3. Where is the hottest part of the burner flame?
4. How would you position a bunsen burner beneath a beaker of water to get the maximum heating effect?
5. Why do you think that it is important to light the match before you turn on the gas tap?

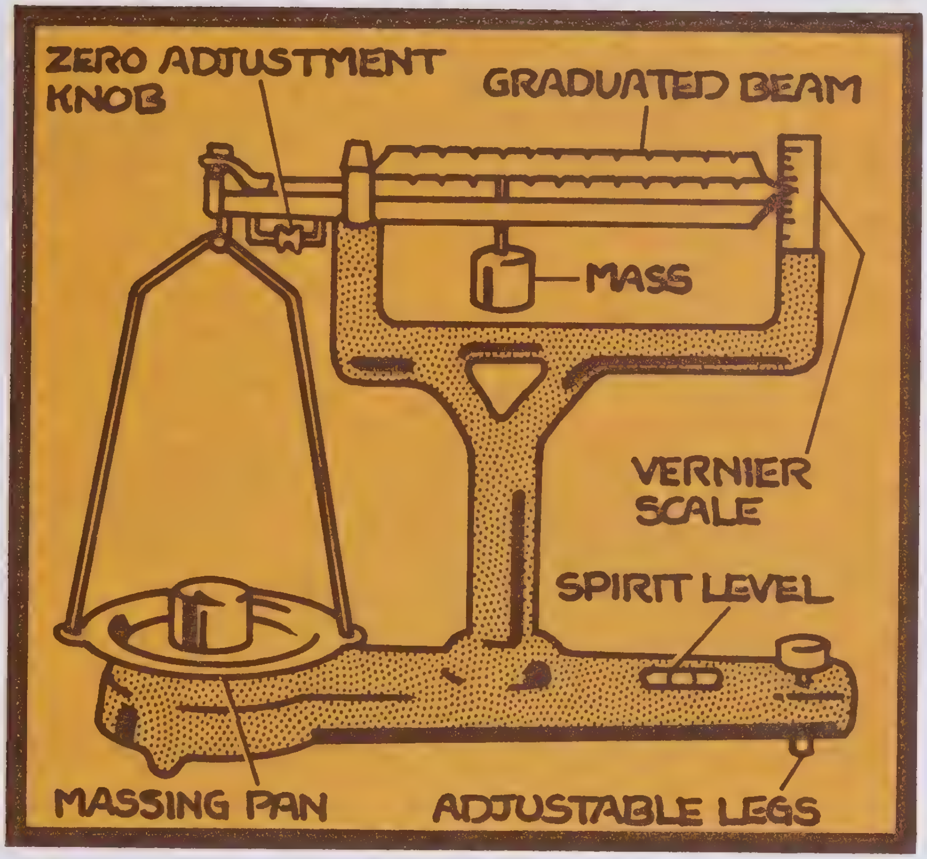
Working with the Balance

Many of the experiments done in chemistry involve determining the mass of something. But unlike the commercial world most of these masses are of very small quantities that must be measured to tenths, or hundredths or even thousandths of a gram. It is surely apparent then that the instrument that you will use must be sensitive. It must be constructed of delicate parts that require a great deal of care and cautious handling.

The balance in the diagram on this page is representative of most of the balances used in chemistry laboratories.

All balances are constructed on a heavy metal base equipped with a spirit level. Adjustable legs permit the balance to be levelled before the mass of anything can be determined. Supported above this base is a beam (or beams in many cases). The beam pivots on hard metal wedges that must be carefully protected for maximum sensitivity.

On one side of the beam is a pan for the object whose mass is being determined; on the other side are movable masses that can be moved along the beams to balance exactly the object on the pan.


ZERO ADJUSTMENT KNOB
GRADUATED BEAM
MASS
VERNIER SCALE
SPIRIT LEVEL
MASSING PAN
ADJUSTABLE LEGS

Before the mass of anything can be determined, all masses on the beams are returned to zero with the pan empty. A knob called the zero adjustment knob is moved along the threaded rod until the beam is balanced; that is, until the mark at the end of the beam, or in some cases a pointer, reads zero on the small vernier scale.

8.2 Determining Masses

balance
10 ml graduated cylinder

Part A

1. Obtain the balance that you will use for this exercise. See if it is equipped with a spirit level and the adjustable legs. Identify the beam or beams. Find the movable masses on the beams and note the graduations on each beam. What are the units?
2. Probably you will find that there are only certain fixed positions that the larger masses can occupy while the smallest mass can occupy any number of different positions.

Determine the maximum capacity of your balance.
Locate the zero adjustment knob.

3. Some balances are equipped with a locking device that prevents the beam from pivoting while not in use. See if your balance has such a device. If so, unlock the balance and lock it again.

Part B

1. Prepare your balance. Let us check. Did you level the balance, unlock it if necessary, return all masses to zero and balance the beam with the zero adjustment knob?
2. Carefully determine the mass of a small graduated cylinder. Record the mass. Have your lab partner determine the mass of the cylinder again. Remember, all masses should have been returned to zero and the balance beam rebalanced before your partner tried. Are the masses the same? If they are not how do you explain any differences?
3. Add 1 ml of water to the cylinder. Determine the mass of the cylinder and water.
4. Repeat Step 3 using 2 ml, 4 ml, 6 ml, 8 ml and 10 ml of water.
5. Calculate the mass of the water in the cylinder in each case. Plot volume versus mass on a graph. Have your lab partner do the same.

Questions

1. Discuss the accuracy of each measurement and the degree of uncertainty that accompanies each measurement.
2. Was the mass of the water in each case that you determined exactly the same as that determined by your lab partner? How do you account for any differences?
3. The graph resulting from plotting mass of water versus volume of water should have been a straight line. Why didn't all points fall precisely on the line?
4. In point form only, summarize the important steps involved in determining the mass of an object.

Cleaning Glassware

Some say that "Cleanliness is next to Godliness." A chemist might say that Godliness comes second in his lab! Even the simplest experiment can be ruined by dirty glassware. While it may sound unnecessary to instruct you how to wash glassware there are a couple of important steps you may not know about.

1. Wash the glassware with warm water and a detergent. Use a brush to clean inside flasks or narrow-necked vessels.
2. Rinse the glassware well in cold tap water.
3. Rinse the inside once with a small quantity of distilled water.
4. Do *not* dry the inside with a towel of any kind.

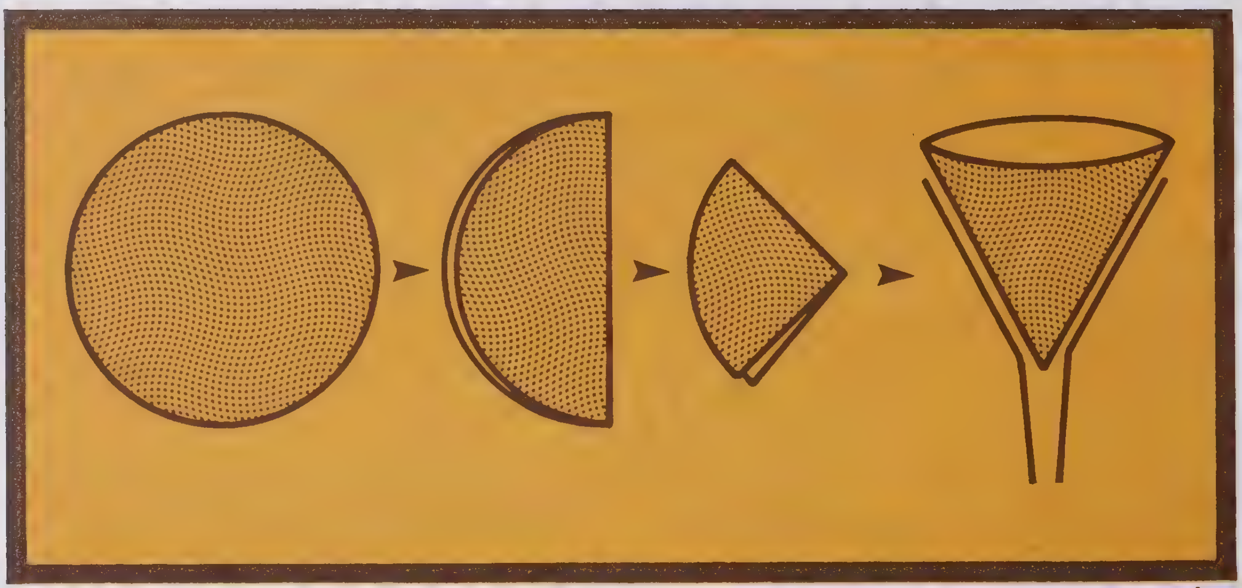
Always clean glassware as soon as you are finished with it. The longer you wait the harder it is to get it clean.

When cleaning burettes the final step is to put a *very* light coating of grease on the stop cock.

Filtration

One of the most frequently used laboratory skills is filtration. Filtration is used to separate a solid from a liquid. The solid is often referred to as a *precipitate*. The principle of filtration is very simple. The solid-liquid mixture is poured onto a porous material (usually paper). The liquid soaks through the paper while the solid is collected on the paper. Special types of filter paper have been made with large, medium or small pores. These enable us to filter different types of precipitates.

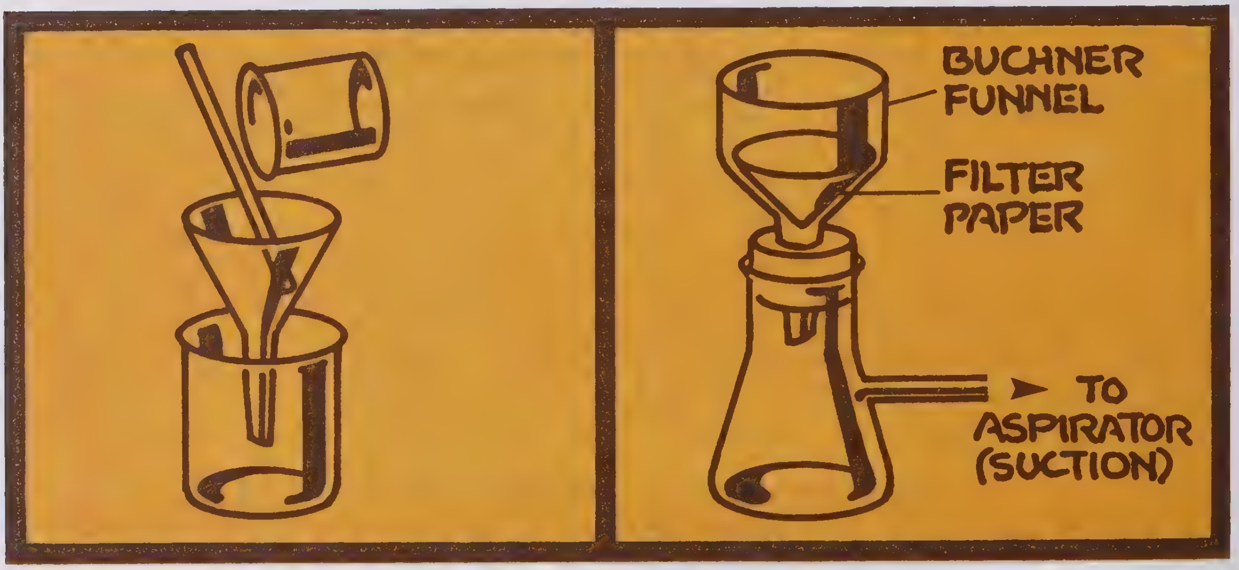
The simplest type of filtration set up is shown in the diagram. The filter paper is folded in quarters then inserted into the funnel with



three thicknesses of paper on one side and one thickness on the other.

A few millilitres of water will keep the filter in position in the funnel.

The mixture to be filtered is carefully poured into the filter cone, using a glass rod to avoid splashing. Normally the first material to run out of the beaker is mostly liquid, while the solid tends to remain until the last. This process is called *decanting*.



Be sure not to over fill the filter cone.

Near the end of the decanting stage it may be necessary to use a wash bottle with distilled water to flush out the last remaining particles of solid. Use as little wash water as necessary.

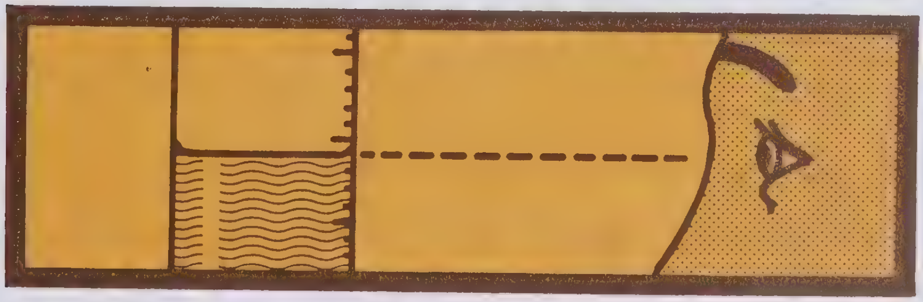
A suction is sometimes used to speed up the filtration process. In this process a *buchner funnel* sometimes replaces the conical funnel.

Measuring Volumes

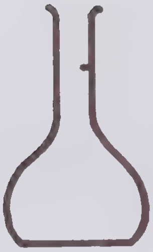
Liquids

In almost every experiment you perform this year you will be required to measure the volume of a liquid. This is normally done with *calibrated* or *volumetric* glassware. Graduated cylinders, volumetric flasks, pipettes and burettes fall into this category. For approximate volume measurements, some beakers are calibrated.

When using volumetric apparatus you must be very careful always to measure the liquid level with your eye level with the *meniscus* or curved surface of the liquid. Not being level with the meniscus causes an error in measurement due to *parallax*.

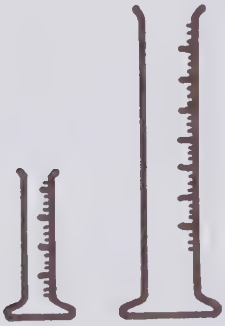


Types of Volumetric Glassware and Their Use



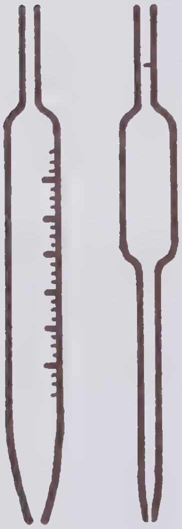
1. Volumetric Flask

These long narrow-necked flasks are calibrated to measure just *one* volume. The flask is filled until the meniscus just rests on the calibration mark.



2. Graduated Cylinders

These tall calibrated tubes can be used to measure any volume. They are most accurate when nearly full. Choose a graduated cylinder which is just large enough to measure the quantity of liquid you need. Don't use a 1000 ml cylinder to measure 8 ml of liquid. Use a 10 ml cylinder. Remember to read the level of the *bottom* of the meniscus on most liquids. (Look at the meniscus on mercury sometime and see why we use the top of a mercury meniscus.)



3. Pipettes

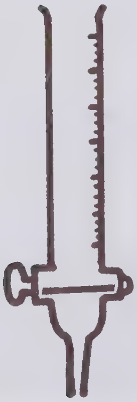
Pipettes come in two types—transfer and mohr. They are used to measure small quantities of liquids to a high accuracy.

Both types of pipette are filled in the following manner. Slip a rubber bulb over the upper end of the pipette. Squeeze some air out of the bulb. Lower the pipette into the liquid and release the bulb. The liquid will rise up the pipette. When the liquid has risen above the desired height slip the bulb off the pipette and immediately cover the open end of the pipette with your index finger tip. (Some people like to use their thumb.)

Very slowly allow the liquid to drain from the pipette until it reaches the desired level. A firm pressure will hold the liquid at any level. Lift the pipette from the solution, blot any excess from the outside with filter paper then transfer the contents of the pipette to whatever container you are using.

A tranfer pipette will deliver only the volume indicated on the pipette. Touch the tip of the pipette to the inside of the receiver to remove the last drop. *Do not blow the last traces of liquid out.*

A mohr pipette will deliver any volume up to the maximum indicated on the pipette. Remember to read the bottom of the meniscus. *Never use oral suction to fill a pipette.*

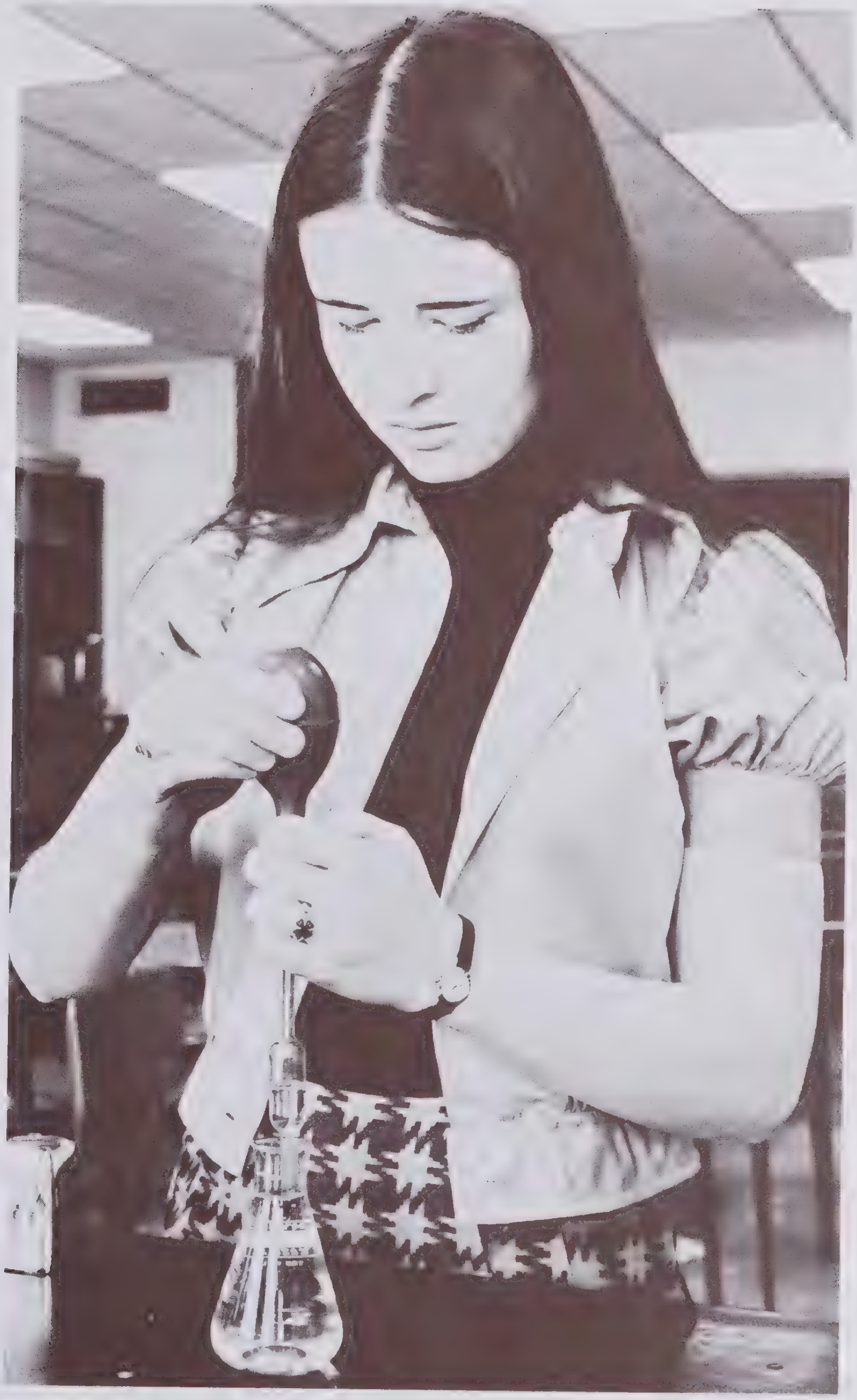


4. Burettes

Burettes are used to measure accurately large or small quantities of a liquid with the added convenience of a tap or stop cock.

Fill the burette past the zero mark using a funnel. Remove the funnel and slowly lower the level until the meniscus rests on zero. Remove the drop left hanging on the tip with filter paper. When using the burette the liquid should be delivered a drop at a time, not in a rapid stream.

One drop of liquid from a burette has a volume of approximately 0.05 ml but with practice a good operator can split drops into halves or even quarters.



To operate the burette hold the body of the stop cock firmly in one hand. Use the other hand to rotate the valve very carefully until the desired drop rate is achieved. If the stop cock is hard to turn it may require a light coat of stop cock grease to free it.

Glass Tubing Manipulations

Glass tubes, ranging in diameter from about 0.5 mm to 15 mm have a variety of uses in a chemistry lab. You should be familiar with several aspects of handling and using glass tubing.

Glassware generally comes in two types, soft (or lime) glass and hard glass such as Pyrex or Hysil.

Cutting Glass Tubing

Tubing up to about 10 mm inner diameter can be "cut" to any length in the following manner.

1. Lay the tubing on a flat surface. Hold it firmly with one hand and draw the edge of a three-cornered file across the tubing at the point where you wish to cut it. This should leave a scratch on the tubing.
2. Wear gloves for this step or wrap the glass in a piece of cloth. Hold the tube with the scratch away from you, one thumb on each side of the scratch. Bend the tubing towards you. The glass should break fairly cleanly at the scratch.

Lot #
Conces
Townsh
Vic

Do not use the glass until it has been fire polished (see next section).

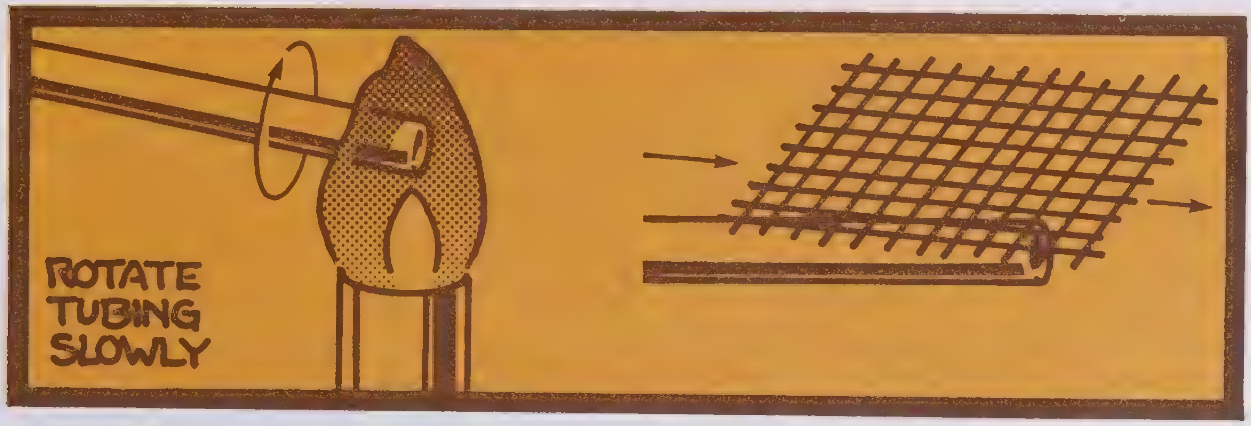
Larger pieces of glass can be cut using a bottle cutter, like those in hobby kits, or by using a hot wire cutter such as the one described in *The Sourcebook for the Physical Sciences.*

Fire Polishing

Freshly cut glass has a very sharp edge. This edge can be removed by rotating the tubing in a burner flame until it melts slightly and the sharp edge disappears. This technique is known as fire polishing.

Larger pieces of glass can have the edge removed with coarse sandpaper or by a metal screen material.

Run the screen over the edge at an angle. Finish with sandpaper.



Warning

If the tubing is held in the flame too long, the opening will become smaller and smaller and eventually disappear.

Inserting Tubing into a Stopper

A very common laboratory accident results when tubing is inserted into a stopper the wrong way and the tubing breaks. Usually one only makes this mistake once.

1. Hold the stopper firmly inside a thick towel with the fingers of one hand.
2. Coat the lower end of the tubing with stop cock grease or vaseline — even water will do.
3. Hold the tubing with a thick towel and twist it gently as it is forced through the stopper. Be *very* careful not to bend the tubing as you push it.

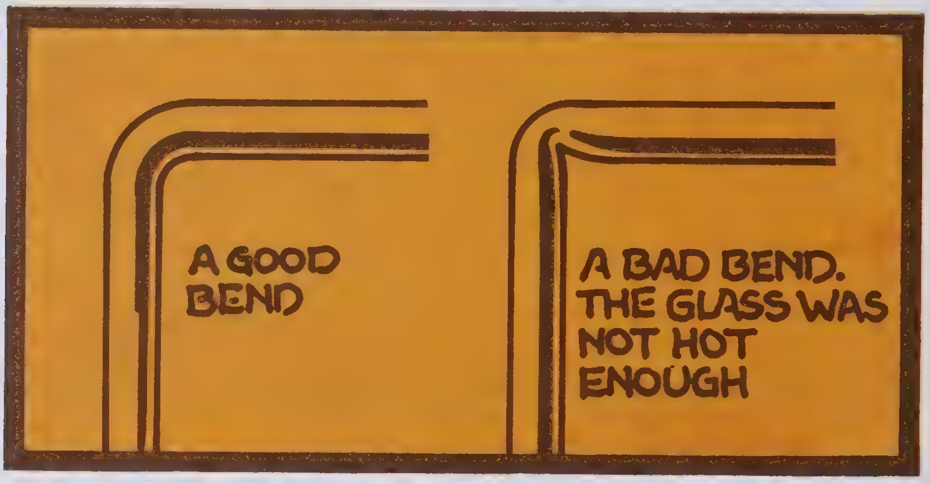
Gloves may be used instead of a towel. Sometimes a cork bore can be forced through the hole first then the tubing inserted into the opening and then the cork bore removed. In any case, regardless of what method is used, exercise a great deal of caution!

Bending Glass Tubing

Pyrex glass cannot be worked in a bunsen flame.

This technique requires a flame spreader or wing top that fits over the barrel of the burner. Always start with a piece of glass that has been fire polished at both ends. You must also protect the desk top with an asbestos pad.

1. Heat the section of tubing to be bent in the flame produced by the wing top. Rotate the tubing constantly.
2. As the glass becomes hot it will become progressively softer. When it has become very soft, remove the glass from the flame and let go of one end. The glass will bend under its own weight.


A GOOD BEND
A BAD BEND. THE GLASS WAS NOT HOT ENOUGH

Making a Capillary Tube

Proceed as for bending tubing until the glass becomes very soft. At this point remove the glass from the flame and, holding firmly to both ends, spread your arms apart. This will draw the glass into a very fine tube called a capillary tube. With practice you should be able to produce tubing of about 1 mm diameter.

Cut a section of this 1 mm tubing about 5 cm long. Hold one end in the flame until it has melted over and has sealed the end shut. This small test tube is called a melting point tube.

Handling Liquids

Since so many of the liquids used in a chemistry laboratory are corrosive, special techniques for handling them have been developed.

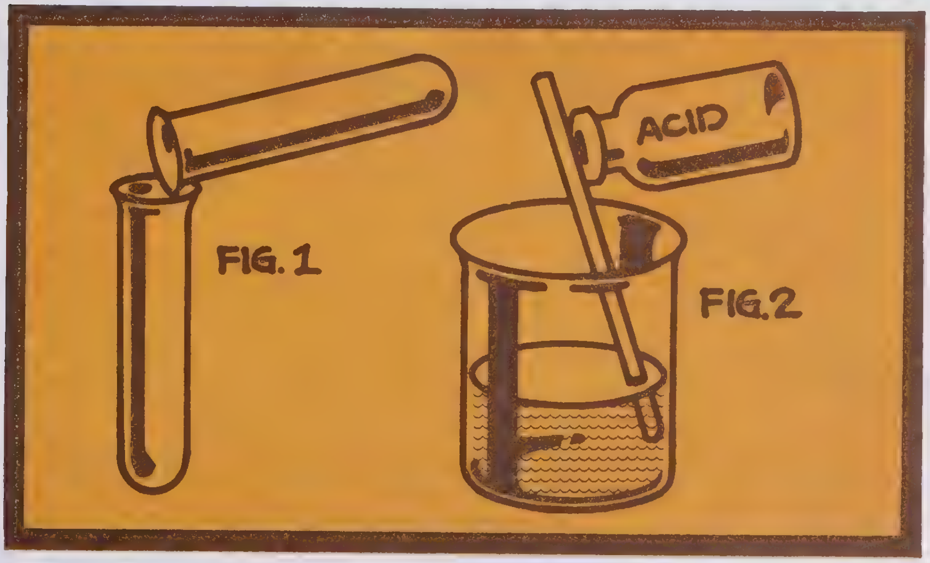
Pouring from Test Tube to Test Tube

This procedure can be done very safely if the following rules are observed.

1. Pour at arms length.
2. Pour over a sink.
3. Hold the lower test tube with the fingers and *not* the palm of the hand.

Pouring into a Beaker or Large Container

When pouring into a beaker pour down a glass stirring rod to avoid splashing (see diagram).



Decanting

Follow the same procedure for pouring into a beaker.

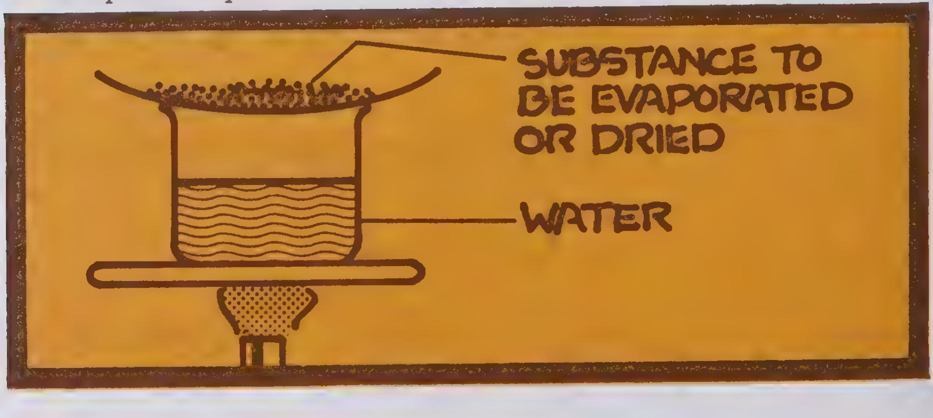
Boiling in a Test Tube

1. Be sure to clamp the test tube at the top.
2. Hold the test tube at an angle of about 45°.

3. Move the test tube gently in and out of the flame to control the rate at which it is being heated.
4. Don't point the test tube at anyone.

Evaporating or Drying over Boiling Water

1. Place the sample to be dried in a watch glass.
2. Place the watch glass on top of a beaker as shown in the diagram.
3. Bring the water in the beaker to a boil. Continue boiling until the sample is dry.

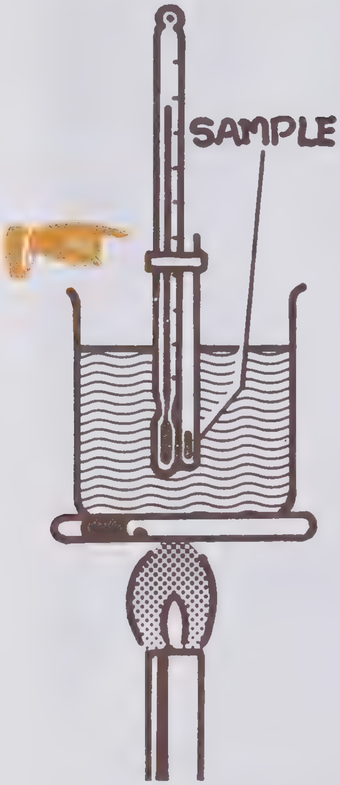

SUBSTANCE TO BE EVAPORATED OR DRIED
WATER

Identifying Odours

In making observations we all recognize that there are five senses available to us. In most cases taste is not used because of the danger of poison. The sense of smell can be used, particularly for gases and volatile liquids, if one does it properly. Here is how to check on odour correctly.

1. Take three deep breaths. Hold the last one in your lungs.
2. Close your mouth and take a gentle sniff of the substance to be checked by wafting it gently toward your nose with your cupped hand.
3. Exhale through your nose.

This last step is most important as it flushes any traces of the gas from your nasal passage.

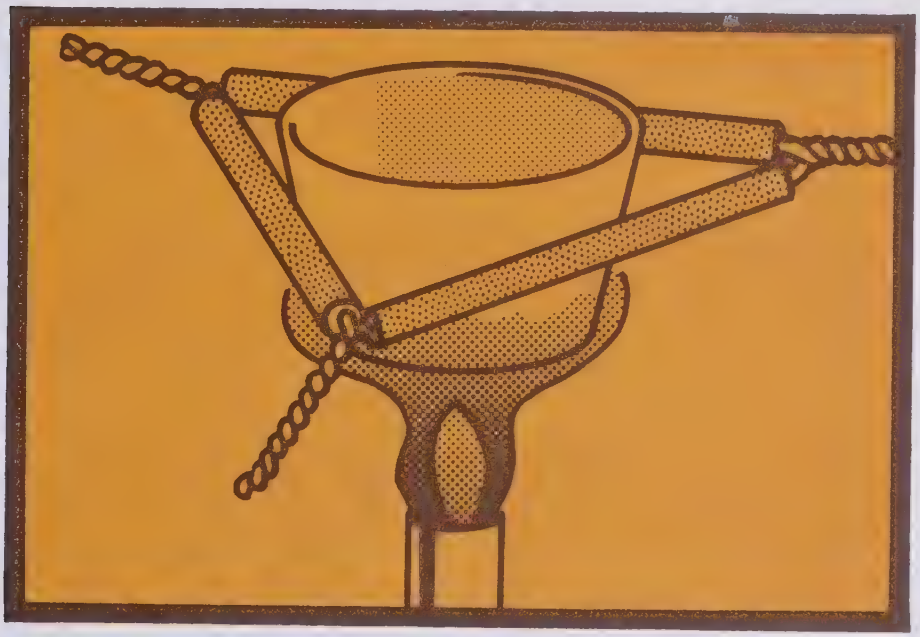


Melting Point Determination Using a Capillary Tube

1. Prepare a capillary tube as in Section 6. Seal one end of the tube.
2. Grind the sample to be tested with a mortar and pestle.
3. Carefully introduce some of the sample into the capillary by inserting the open end into the sample and then tapping the tube gently on a firm surface to force the sample to the bottom of the tube. You usually only need 5 mm of the sample.
4. Attach the capillary to a thermometer with a small elastic band cut from rubber tubing. Lower into a beaker of water.
5. Adjust the sample so it is opposite the bulb of the thermometer, but do not immerse the open end of the capillary below the surface of the liquid.
6. Heat the liquid slowly. Record the temperature at which the sample melts.
7. Repeat at least twice more.
8. For samples that melt at temperatures above 100°C substitute mineral oil or cooking oil for water as the heating liquid.

Heating in a Crucible

A well-adjusted bunsen burner flame reaches a maximum temperature of nearly 1000°C. It is possible, with some care, to use this flame to heat small quantities of a substance up to a temperature of about 800°C. This requires a special vessel called a *crucible* which is made of a ceramic material and is very resistant to high temperatures. The critical factor in achieving a high temperature inside the crucible is its position in the flame.



1. Select a clay triangle that holds the crucible about half way up the sides.
2. Light and adjust the burner for maximum heat (see Exercise 1).
3. Clamp the triangle in position so that the bottom of the crucible is about 3 cm above the top of the inner blue cone.
4. A well-fired crucible should glow a bright cherry red colour while it is being heated. You may have to adjust the height slightly to achieve this colour.

Warning

Crucibles that have been heated remain hot for a *long* time!

9

Food Chemistry/
Molecular Masses/
Industrial Processes/
Molecular Size/
Crystals/
Chromatography/
Pollution/Dyes/
Electrolysis

Introduction

This final unit contains some topics that are suggested as independent study projects. Each student or a group of students, preferably no more than two, is expected to select a topic for independent study. You may find one of the topics listed here will interest you. If not, your teacher will probably listen to your suggestions for a topic and may give you permission to do some topic not listed here.

A project should be carried out in two phases:

1. preparation;
2. execution.

The second phase—execution—will be carried out in the laboratory. Your teacher will set aside a number of days for this phase. You will work at your project in the laboratory during those periods and complete the practical part of your project.

The first phase — preparation — will be carried out at intervals over an extended period. Your teacher will outline the steps required of you and when each step must be completed so that you will be ready to start the laboratory phase.

In general terms, you should try to do some reading around your topic so as to get a general idea of what it is about. You should try to work up some experiments that you might perform in the laboratory and prepare a list of materials for each of them. You will then be able to discuss with your teacher how to get those items and you should decide in advance how you intend to present the completed topic to your teacher. Students have presented projects like these in a variety of ways. Many of course choose the traditional written report but others have made motion pictures, produced a photographic report, a demonstration of an experiment, a lecture to the class, a tape recording, a series of drawings or diagrams and many combinations of the above. There can be no doubt that the more original you are the more impressive your project will be.

Whatever else you do you must try to convince your teacher of at least two things:

1. that you did something;
2. that you learned something.

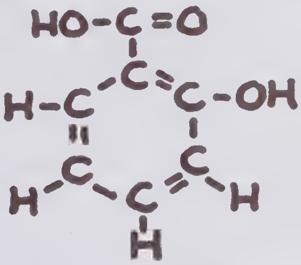
This is your opportunity to do some work on your own and to demonstrate what you can do if you are left alone. It should be the best part of the course for you.

Drugs

One of the most common drugs on the drugstore shelf is one that chemist call *acetasalicylic acid*. If you take the most prominent letters out of that name you have ASA. This white powder is pressed into tablets and marketed under the trade names of Aspirin, Anacin, ASA tablets and many others. It is also sold in combination with other drugs in pain killers, cold preparations, etc.

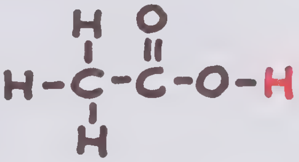
The preparation of ASA is fairly straightforward.

Salicylic acid has the formula



The replaceable hydrogen or the acid hydrogen is the one we have shown in red. The OH shown in brown is called an alcohol group.

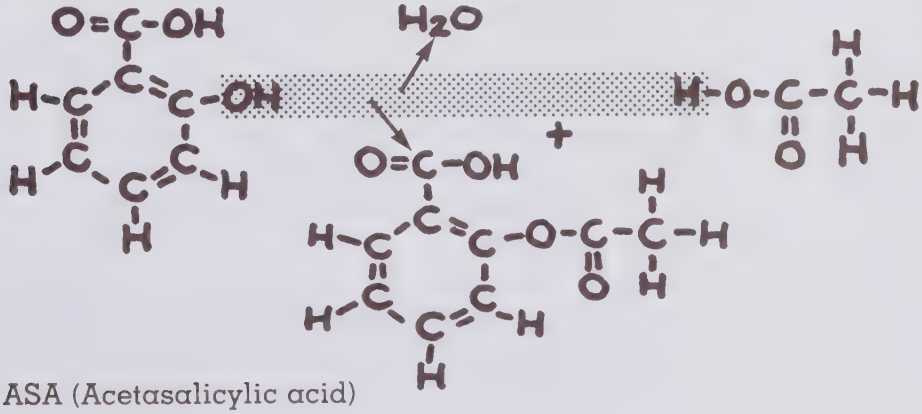
Acetic acid has the formula



$$\mathrm{H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle O}{\overset{\|}{C}}-O-H}$$

Again, the acid hydrogen or the replaceable hydrogen is shown in red.

These two substances can be made to combine in a reaction similar to neutralization. The replaceable hydrogen from acetic acid reacts with and removes the alcohol group from the salicylic acid. The two pieces that are left then join together to form ASA.



ASA (Acetasalicylic acid)

salicylic acid
acetic anhydride
18 M sulphuric acid
ice
suction filter
125 ml erlenmeyer flask

1. Place 5 g salicylic acid in a 125 ml erlenmeyer flask.
2. Add 10 ml acetic anhydride.
3. Add 8 to 10 drops of 18 M sulphuric acid.

Warning

18 M H_2SO_4 is very caustic.

4. Swirl the ingredients gently for 3 to 5 min. Allow the flask to stand at 35-40°C for 8 to 10 min.
5. Cool the flask in an ice bath to below 5°C.
6. Add 100 ml ice water. Mix well.
7. Filter with suction (see Unit 8, Exercise 4).
8. Wash the precipitate with ice water.
9. Dry on a watch glass.
10. Determine the mass of your ASA product.
11. Determine the melting point of your ASA product using a capillary tube (see Unit 8, Exercise 9).

Warning

Under *no* circumstances should you sample the product. The deviation between your melting point and the accepted melting point of 135°C is a measure of how impure your sample is.

Suggested Experiments

1. Isolation of caffeine from tea or coffee.

Sources of Further Information

1. *Tested Demonstrations in Chemistry*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from the *Journal of Chemical Education*, American Chemical Society, Easton, 1965.
2. Any organic chemistry text.

Food Chemistry

We all differ from each other in many ways but among the few things we all have in common is the need to eat. We all need food to fuel our internal fire. Different foods contain different substances. Our bodies convert these different materials into various essentials; energy, tissue, blood, fat and bone, to name just a few. In this project you will find the procedures for testing for various nutrients in food. All you need supply is the actual food.

Some Simple Nutrient Tests

Part A Glucose

1. Place a sample of the food to be tested in a test tube. Add 5 ml of water and 5 ml of Benedict's solution (or Fehling Solution). Warm the contents gently in a water bath.
2. Record any noticeable change in colour.

Part B Starch

1. Place a sample of the food to be tested in a test tube. Add 5 ml of water and 5 ml of Lugol solution.
2. Record any noticeable change in colour.

Part C Fats

1. Rub a sample of the food to be tested on a piece of unglazed paper, such as a paper towel. Hold the paper up to the light from a window or fluorescent fixture.
2. Record any marks that have appeared on the paper.
3. If the fats are suspected but *not* indicated by this test, then do the next test. This should be carried out in a fume hood. Grind the food to be tested with a mortar and pestle with carbon tetrachloride. When the grinding is complete, pour some of the liquid onto the paper towel and repeat Steps 1 and 2.

Warning

Carbon tetrachloride fumes are very dangerous. This should *only* be done in a well-ventilated area.

Part D Protein

1. Place a sample of the food to be tested in a test tube. Just cover the sample with dilute nitric acid. Warm the test tube in a water bath.

2. Carefully decant the liquid and record any colour changes that have occurred.
3. Cover the sample with dilute ammonium hydroxide. Record any colour changes that occurred.

Alternate Protein Method

1. Place a sample of the food to be tested in a test tube. Add 5 ml of distilled water. Boil the test tube for 3 to 5 min.
2. Add two drops of (15 percent in alcohol) β-napthol solution.
3. Incline the test tube sharply and carefully pour 3 ml concentrated sulphuric down the side so as to form two layers (as in the brown ring test for nitrates). Note any colour formed at the interface.

Part E Water

1. Heat a sample of the food to be tested in a sharply inclined test tube. Test any condensate around the mouth of the test tube with cobalt chloride paper or anhydrous copper sulphate.

Part F Ash or Mineral

2. Heat a sample of the food to be tested in a clean crucible using the full heat of the burner for about 15 min. Record any observations on any residue.

Suggested Experiments

1. Bring samples of food from home for testing.
2. Test a complete meal for balance.
3. Compare your analysis with that of the Canada Food Guide.
4. Warm some table sugar with a few drops of hydrochloric acid for 10 min to test for glucose.
5. Mix a sample of starch with saliva and then test for glucose.
6. Compare the sweetness of various sugars and sugar substitutes.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Sourcebook for the Biological Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World.
3. *Canada Food Guide*. Ottawa: Queen's Printer.

Nutrients in Milk

Have you seen the television commercial that says, "Drink milk for all it's worth"?

Think about it for a moment and that phrase can take on a lot of different meanings. One meaning could be that there is a lot of nutrition in milk. Here is a project that will let you determine just what it is in milk that makes mothers the world over say, "Drink your milk, dear."

To do this project you need to be familiar with nutrient tests such as those used in the previous project.

Some Tests for Nutrients in Milk

milk
500 ml beaker
distilled water
0.25 M acetic acid
filter paper
funnel
0.1 M NaOH

Part A

1. Pour 25 ml of milk into a 500 ml beaker. Add 175 ml distilled water. Add slowly, drop by drop with constant stirring, 40 ml of 0.25 M acetic acid. Allow the solution to stand for 20 min. Filter and save both the residue and the filtrate. (Proceed with Part B while you are waiting for the precipitate to settle.)
2. Rinse the precipitate from the filter into a beaker. This material is called *casein*. Test a small quantity of it for sugar, protein, starch and fat.
3. Take the remaining casein and pour off as much liquid as possible. Slowly add 100 ml of 0.1 M NaOH with constant stirring. If the casein does not dissolve add more 0.1 M NaOH slowly.

Questions

1. What causes milk to curdle when it enters the stomach?
2. How could casein be reprecipitated from the NaOH solution?
3. Give a rough estimate of the amount of casein in milk.

Part B

250 ml beaker
filter paper
milk
watch glass
carbon tetrachloride
bunsen burner
ring stand

1. Half fill a 250 ml beaker with water. Heat the water to boiling on a ring stand.
2. While waiting for the water to boil, wet a 6 cm circle of filter paper with fresh whole milk. Place the soaked filter paper on a watch glass on top of the boiling water and allow the filter paper to dry. Remove the paper and watch glass and allow them to cool.

Warning

This next section must be performed in the fume hood using rubber or plastic gloves. Carbon tetrachloride fumes are dangerous.

3. Pour sufficient carbon tetrachloride into the watch glass to cover the filter paper. Work the paper around in the liquid for 2 to 3 min using a stirring rod. Remove the paper and evaporate the carbon tetrachloride over a hot plate.

Questions

1. What remains on the watch glass?
2. What is carbon tetrachloride commonly used for?

Part C

filtrate from Part A
bunsen burner

1. Heat the filtrate obtained in Part A until a clear fluid can be obtained by filtration. Save the clear filtrate for Part D. The residue on the filter paper is called *lactalbumen*. Test this material for protein, starch and fat.

Questions

1. What type of material is lactalbumen?

2. Casein does not coagulate when milk is boiled. What is the scum that appears on the surface of boiling milk?

Part D

filtrate from Part C

1. Take half a test tube of the clear filtrate obtained in Part C and apply the test for sugars to it.

Questions

1. The sugar found in milk is called *lactose* or *milk sugar*. Taste a sample provided by your teacher. How does it compare to sucrose or cane sugar?
2. Certain bacteria can convert lactose into lactic acid. What will happen to milk when this takes place?

Part E

milk
bunsen burner
crucible

1. Evaporate 5 ml of whole milk to dryness in a crucible. Then fire the crucible under the full heat of a burner for about 5 min.

Questions

1. What type of nutrients are identified by this procedure?
2. List some of the nutrients that this series of experiments has identified as being present in milk.
3. Are you aware of any other nutrients present in milk?
4. Now do you know why your parents are always after you to "Drink your milk"?

Molecular Masses

The developement of Avogadro's Hypothesis led directly to a means of determining the relative masses of molecules; that is, their molecular masses. All one has to do is find the mass of equal volumes of two gases measured at the same temperature and pressure. Since the two volumes contain equal numbers of molecules then a ratio of their masses gives their relative masses.

Understand what we mean here by relative masses. Suppose you wanted to know the relative masses of baseballs and tennis balls, but the only scale available is graduated in kilograms. What do you do? Rather than attempting to find the mass of just one of each, you find the mass of 50 of each. The ratio of the mass of 50 baseballs to the mass of 50 tennis balls is exactly the same as the ratio of the mass of one baseball to one tennis ball.

With gaseous substances the mass of equal volumes is found and a ratio of their masses is equal to a ratio of the masses of individual molecules. All molecular masses appear on the same scale and further, on the same scale as atomic masses. Hence, by agreement, the molecular mass of oxygen is 32.

The reference standard for molecular masses is oxygen = 32.

The molecular mass of a substance is a number representing the mass of the molecules of that substance compared to the mass of an oxygen molecule taken as exactly 32.

Molecular masses may be determined fairly easily, especially for gaseous substances or substances that can easily be converted to the gaseous state.

Determination of the Molecular Masses of Several Gases

oxygen gas
carbon dioxide gas
tap gas
jug

1. Find the mass of equal volumes of different gases in the same manner as you did in Experiment 2.4. One of the gases used should be oxygen.
2. Determine the volume of the container by the downward displacement of water.
3. Determine room temperature and the density of air at that temperature. Calculate the mass of air in the container used above.
4. Calculate the actual mass of each of the gases used in Step 1 (apparent mass plus mass of air in the container).
5. For each gas used in Step 1 determine the ratio: mass of gas/mass of equal volume of oxygen.
6. If oxygen gas is given a molecular mass of 32, calculate the molecular mass of each of the other gases.
7. Compare the molecular mass from Step 6 to the actual molecular mass for each gas as found in the *Handbook of Chemistry*.

Suggested Experiments

1. Determine the molecular mass of a volatile liquid by vapour pressure.
2. Illustrate the concept of relative masses using seeds.

Industrial Processes

Physical and chemical laws underlie all the industrial processes in use in the world today. Sometimes the laws are clear and well defined, sometimes not. The term industrial process is generally used to refer to any process — physical or chemical — that transforms raw materials into finished products. Processes involving chemical reaction are of primary interest in this project.

The student should select some industrial process — preferably one nearby that he can visit—and attempt to outline this process in terms of the physical and chemical laws, principles and changes that occur during the process.

For example, a sawmill is an industrial process. Trees — the raw material — are converted to boards — the finished product — in a completely physical process. It should not interest you for this project because no chemical processes take place. Be sure the process you select contains a chemical reaction.

The list of possible processes is endless. You might examine:

1. oil refining
2. paper making
3. smelting of ores
4. manufacture of chemicals
5. distillation and fermentation
6. manufacture of drugs
7. manufacture of paints
8. manufacture of soaps
9. manufacture of cosmetics

Because of the extremely broad nature of the topic no sample experiment can be included. Your project should take the following steps:

1. selection of the process;
2. determination of the flow diagram either from library research or a visit to the plant;
3. analysis of the process in terms of the physical and chemical principles involved;
4. (optional) demonstration of some phase of the process in the laboratory.

Sources of Further Information

1. A letter to the company involved.
2. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
3. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
4. *Chemical Process Industries*. Shreve. New York: McGraw-Hill, 1967.
5. Any encyclopedia.
6. *Principles of Unit Operations*. Foust et al. London: Wiley, 1960.

Cosmetic Chemistry

What needs to be said about cosmetics? Most women in the Western world use something in the cosmetic line at one time or another. Cosmetics are the things that people apply to their bodies to cleanse or beautify them. Although soap fits such a definition it is not normally included as a cosmetic.

Cosmetics have been around for about 6000 years. If you think that eyeshadow is a recent inovation, remember Cleopatra used it.

The list of preparations is endless. We have listed three very good sources of material for you to choose from. For your introductory experiment here is a recipe for a cold cream.

Making Cold Cream

white beeswax
2 150 ml beakers
liquid paraffin
bunsen burner
sodium borate
perfume

1. Put 15 g of white beeswax in a 150 ml beaker. Add 65 ml of liquid paraffin. Warm this mixture to 80°C. Mix the now molten beeswax with the liquid paraffin.
2. In a second 150 ml beaker add 1 g of sodium borate (Borax) to 30 ml of water. Warm this mixture to 80°C with stirring to dissolve the sodium borate.

3. Carefully and with constant stirring pour the second mixture into the first. When all the mixture has been added, remove the mixture from the heat but continue stirring vigorously until it has cooled to 30°C. Add a few drops of perfume and again stir well. Cool to room temperature. Transfer to a container with a lid.

Suggested Experiments

1. Prepare nail polish.
2. Prepare lipstick.
3. Prepare toothpaste.
4. Use your imagination.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *The Chemical Formulary*. Bennett, H. New York: Chemical Publishing Co., 1945.
3. *Cosmetic Science*. Griffin Technical Studies. Middlesex, England: Griffin and George, 1971.

Measuring Molecular Size

In Unit 2 we tried to estimate the size of an atom. Here is a way of getting an approximate size for a molecule. The method was first recorded by Dr. Irving Langmuir in 1917. It is regarded by many as a classic experiment.

The molecule Dr. Langmuir chose to work with was oleic acid ($C_{17}H_{33}COOH$).

It looks like a large molecule. We might even assume that it will be larger than an atom.

The Size of a Molecule of Oleic Acid

dilute oleic acid
ethyl alcohol
beaker
10 ml graduated cylinder
large tray
lycopodium powder

1. Prepare a dilute solution of oleic acid in alcohol by first dissolving 5 ml of oleic acid in 95 ml of ethyl alcohol. This solution will be 5 percent oleic acid.
2. Dissolve 5 ml of the 5 percent oleic acid solution in a further 45 ml of ethyl alcohol. This solution will be 0.5 percent of oleic acid.
3. Count the number of drops of this 0.5 percent solution required to fill a 10 ml graduated cylinder.
4. Obtain a large tray (such as a ripple tank used in the physics lab). Pour in about 2 cm of water. Allow the water to come to room temperature.
5. Lightly sprinkle lycopodium powder or chalk dust over the surface of the water.
6. Add one drop of the 0.5 percent oleic acid to the centre of the tray. Measure the dimensions of the pattern that forms.

Calculations

1. Given that the density of oleic acid is 0.9 g/ml determine the mass of oleic acid in a single drop of 0.5 percent oleic acid solution.
2. Calculate the volume of oleic acid in a single drop of 0.5 percent oleic acid solution.
3. Knowing that the pattern formed on the surface was formed by a thin layer of oleic acid left behind as the alcohol dissolved in the water, calculate the area covered by the oleic acid.
4. Knowing the volume of the oleic acid in a single drop and the area the drop covered, calculate the thickness of the oil film on the surface of the water.

Questions

1. If the layer of oil is one molecule thick, what is the thickness of a molecule?
2. How does this figure compare to the atomic diameter figure obtained in Unit 2?

Suggested Experiments

1. Repeat this experiment using two drops of 0.5 percent oleic acid instead of one.
2. Repeat using substances such as stearic acid in place of oleic acid. (In this case the tray must contain hot water.)

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.

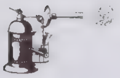
Measuring Inter-Molecular Energies

Molecules are very small. So small that they cannot be seen with the unaided eye. Even with microscopes we cannot view the world of molecules. But what of the forces that hold molecules together? Can these be measured?

This project will introduce you to the way in which one might estimate the so-called *bond forces* by showing you how to estimate the force of attraction between different types of molecules. You should familiarize yourself with the last project on estimating molecular size as we will need that information later on.

If you like physics you should love this sort of project.

Molecular Bonds



wire
metal or plastic lid 5 cm in diameter
beaker
balance

1. Glue or solder a wire to the inside of a metal or plastic lid about 5 cm in diameter.
2. Place an empty beaker with a larger diameter than the lid on the support pan of the balance.
3. Fasten the lid to the balance beam so that it is about half way into the beaker and so the lower surface of the lid is level.
4. Balance the beam so the pointer rides at zero. Record this mass.
5. Carefully add water down the side of the beaker until the lid is *just* touching the surface of the water.
6. Move the riders on the beam until the lid breaks free from the surface of the water. Record this figure.

Calculations

1. Given that the density of water is 1 g/ml and GMM is 18 g, calculate the approximate size of a water molecule. (Assume they are spherical.)
2. Calculate the area of water in contact with the surface of the lid.
3. How many water molecules were in contact with the lid?
4. What force was required to overcome the force of attraction of these water molecules for the lid?
5. Calculate the force exerted by each molecule.

Suggested Experiments

1. Repeat this experiment using a different material for the lid or coat the lid with other substances.
2. Repeat using a different liquid in the beaker.
3. Repeat with various quantities of soap or detergent added to the water.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. Any college level physics book.
3. Encyclopedia. Look under Surface Tension.

Measuring the Strength of a Bond

Here is a project that will help you understand something about how we are able to measure the energy of a bond. As with all experiments of this type, it is almost impossible to measure a single bond energy. Instead we measure the strength of many identical bonds and then divide by the total number. We have selected the hydrogen-oxygen bond in water as the one to try and estimate. Remember that there are two such bonds in a water molecule.

Measuring a Covalent Bond

apparatus for Experiment 3.10
distilled water
sulphuric acid
voltmeter
ammeter

1. Set up the apparatus used for the electrolysis of brine back in Experiment 3.10.
2. Use distilled water with a small amount of sulphuric acid (1 percent) as the liquid.

3. Add a voltmeter and an ammeter to the circuit. The ammeter should be in series and the voltmeter in parallel across the electrodes.
4. Prepare a table to record voltage, amperage and time for the duration of the experiment.
5. When all is ready switch on the power supply and quickly adjust it to whatever setting gives a reasonable supply of bubbles off the two terminals.
6. Record the voltage and amperage every 30 s until you have collected a measureable amount of gas.
7. Record the total time the apparatus was in operation and the total volume of each gas collected.

Calculations

1. From the amount of gas collected calculate the number of water molecules decomposed and then the number of O-H bonds that were broken.
2. From the *average* current and the *total* time, calculate the number of coulombs of electricity used ($Q = A \times t$ (s)).
3. By multiplying the number of coulombs of charge used by the average voltage, determine the number of joules of electrical energy used in this experiment.
4. Assuming that all this energy was used to break bonds, determine the bond energy of one O-H bond.

Suggested Experiments

1. Repeat this experiment using other substances such as HC1.
2. Compare the internal bond energy with an external bond energy.

Sources of Further Information

1. A college-level chemistry test.

Pulp and Paper

Probably no industry sounds more Canadian than pulp and paper. Each year this industry brings in millions of dollars of foreign business and provides thousands of Canadians with jobs. Perhaps the biggest assets of this industry are that it uses a renewable resource — trees — and that it can easily adapt to the pollution combatting process called *recycling*.

The paper making process is basically simple. Trees are composed of long fibres of cellulose held together by a substance called lignin. In special "cookers" lignin is separated from these fibres so they may be rearranged. This rearrangement is done in the form of thin flat sheets that, when dried, are called paper.

This project will show you some aspects of the process.

Making Paper

white newsprint
1000 ml beaker
bunsen burner
wire screen
tightly woven cloth
wax paper
plywood

1. Obtain a supply of white newsprint. If necessary just cut the margins off a newspaper.
2. Cut the paper into 2 cm strips and place them in a 1000 ml beaker until the beaker is half to three-quarters full. Add enough water to cover the paper.
3. Bring the contents to the boil and continue to boil until the mixture is a uniform soupy mass. There must be no lumps.
4. Flatten a piece of wire screen on a desk. Cover it with a piece of tightly woven cloth.
5. Spoon out pulp onto the cloth and spread it so that it is uniformly about 0.3 cm thick. Cover the pulpy layer with a piece of wax paper and then a piece of plywood.
6. Flip the whole apparatus over and begin to press away the moisture with a sponge. When the pulp is as dry as it can be patted, place it in a warming oven and drive off the moisture that remains.
7. Remove the apparatus from the oven, gently peel back the screen and the cloth. The product is very crude paper.

Suggested Experiments

1. Improve the quality of your paper by building the paper machine described in *Science Affairs,* September 1967.
2. Try to make paper from wood instead of recycling paper.
3. Waterproof or fireproof some paper.

Sources of Further Information

1. *Tested Demonstrations.* Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
2. *Sourcebook for the Physical Sciences.* Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
3. *Chemistry—Collected Experiments.* Nuffield Foundation. London: Penguin, 1965.

Indicators

Two of the commonest chemical families are those of the acids and the bases. Rare indeed is the adult who has never heard the word acid. But what is an acid? How do we identify acids and bases?

In nature there are substances such as the juice of a red cabbage or the juice from blueberries that can change colour if they change from an acid solution to a basic solution. Substances that do this are called *indicators* and can be used to identify acids and bases.

They can be used for other things of course; for example, the indicator phenolphthalein is used as a laxative. Indicators can also be used to tell how strong an acid or a base is.

Acid Strength

1 M HCl
30 test tubes
distilled water
1 M NaOH
bromthymol blue
phenolphthalein
orange IV

1. Put 10 ml of 1 M HCl in a test tube. Label it Solution A.
2. Take 1 ml of Solution A and put it into a second test tube. Add 9 ml of distilled water. Label it Solution B.
3. Take 1 ml of Solution B and place it in a third test tube. Add 9 ml of distilled water. Label this Solution C.
4. Take 1 ml of Solution C and place it in a fourth test tube. Add 9 ml of distilled water. Label this Solution D.
5. Take 1 ml of Solution D and place it in a fifth test tube. Add 9 ml of distilled water. Label this Solution E.

6. Divide each solution into three equal parts so that you have three test tubes of Solution A, three Solution B and so on for a total of 15 test tubes of acid solutions.
7. Repeat this procedure starting with 1 M NaOH to produce Solution A_1, B_1, C_1, D_1, E_1. You will now have 15 test tubes of basic solutions.
8. Keeping track of things is the hard part. Now add two drops of the indicator bromthymol blue to one set of Solutions A, B, C, D and E and one set of A_1, B_1, C_1, D_1 and E_1.
9. Add two drops of the indicator phenolphthalein to a second set of solutions and two drops of the indicator orange IV to the third set of solutions.

Suggested Experiments

1. Repeat this experiment using other indicators.
2. Test some common substances such as soap, pickle juice, baking soda, sugar, salt, flour, soft drinks and so on for acids and bases.
3. See if you find a natural indicator in your home.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.

Automobile Chemistry

In their efforts to get more power from smaller engines, auto manufacturers have tried many things. One easy way to increase power is to increase the compression ratio of the engine. Car enthusiasts have planed the cylinder heads of cars for years to increase compression and hence get more power from the same engine.

Unfortunately, higher compression ratios lead to problems. One of these is "pre-ignition" or "knocking." As the air-fuel mixture is compressed it becomes hot. If compressed enough it will ignite spontaneously without the need of a spark plug. This pre-ignition has disastrous effects on the power output of the car.

To overcome this problem petroleum engineers add chemicals to gasoline to prevent or delay pre-ignition. One such additive is tetra-ethyl lead. Perhaps you could begin this project by looking for lead in gasoline.

Lead in Gasoline

filter paper
gasoline
1 M acetic acid
1 M potassium iodide
dropper

Warning

This experiment is potentially dangerous. Do it in a fume cabinet.

1. Soak a circle of filter paper in ordinary gasoline. Do this in a well-ventilated area. There must be no open flames anywhere during this experiment.
2. Allow the filter to dry. Expose the dry filter paper to direct sunlight for 5 min.
3. Using a dropper moisten the filter paper with 1 M acetic acid. Again allow it to dry.
4. Again using a dropper moisten the filter paper with 1 M potassium iodide.
5. The amount of yellow formed is proportional to the amount of lead present in the gasoline.

Suggested Experiments

1. Compare different brands of gasoline for lead content.
2. Compare high octane and low octane gasoline for lead content.

Suggested Areas for Further Experiments

1. Determine the viscosity of lubrication oils and compare your findings to the Society of Automotive Engineers (SAE) ratings.
2. Measure the boiling point, freezing point and heat capacity of anti-freeze and compare it to pure water.
3. Look for the presence of carbon monoxide in automobile exhaust fumes.

Sources of Further Information

1. *Sourcebook of the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.

2. *Nuffield Science Materials*. Toronto: Longmans Canada, 1967.
3. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.

Crystals

Earlier in this book the three states of matter were discussed and differences and similarities were indicated. We said then that solids are rigid. This is because the atoms forming a solid are unable to move past each other. The atoms are in motion, however; this is shown by the fact that solids sublime. Some, like dry ice, do so very rapidly. The atoms must be in motion in order to escape from the solid. The atoms in solids stay in one place on the average and their motion is restricted to vibrations about that one place.

The fixed average positions about which the atoms vibrate are arranged in an orderly fashion. The order itself is in a particular pattern repeated over and over again. This could be compared to a child's building blocks placed in the form of a cube. The cube is just layer upon layer of blocks arranged in a regular fashion.

Orderliness of this regularly repeated sort is called *crystallinity*. Any substance possessing crystallinity is a *crystal*. Crystals are beautiful solids with flat faces and regular angles.

Diamond or quartz are examples of crystals with these characteristics. Many substances on the other hand may not have a sparkly appearance like a diamond yet possess the orderliness of a crystal. Metals for example are crystals. Even small irregular fragments may still be crystals.

Crystals may be prepared with some time and a great deal of patience. Perhaps prepared is not a suitable word here since crystals are actually grown from solutions.

A saturated solution is prepared and then seeded. A seed crystal is a small crystal of the substance in solution. Seeds are made by evaporating a small amount of a saturated solution to dryness. They should be about 0.3 to 0.6 cm in length.

Follow the procedure below and attempt to grow a crystal of your own. If you are successful the results will be very satisfying. The substance to be used is called potassium aluminum sulphate dodecahydrate or just *alum*.

Growing Crystals

alum solution
open jar
silk thread
cheese cloth

1. Prepare a saturated solution of the alum at the same temperature at which the crystal will be grown.
2. Pour about 25 ml of this solution into an open jar and let it stand for a few days. As the water evaporates small crystals will start to form in the bottom. Recover one of these crystals. It will be the seed crystal.

The remainder of your solution should be in a sealed container while you are preparing the seed crystal.

3. Place the saturated solution that you have prepared in an area where it will not be disturbed and also where the temperature will remain fairly constant.
4. Suspend the seed crystal into this solution using a silk thread.
5. Cover the jar with a cheese cloth and leave it untouched until the crystal has grown.

Suggested Experiments

1. Grow crystals of:
   (a) chrome alum
   (b) Rochelles salt
   (c) nickel sulphate.
2. Prepare a "crystal shrub."
3. Prepare a "crystal cloud."
4. Prepare a "crystal crown."

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph A. et al. New York: Harcourt, Brace and World, 1961.
2. *Crystals and Crystal Growing*. Holden and Singer. London: Heineman Educational Books, 1961.
3. *Teaching Chemistry with Models*. Sanderson. Princeton: Van Nostrand, 1962.

4. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
5. *Modern Chemistry*. Metcalfe, H.C. et al. New York: Holt, Rinehart and Winston, 1970.

Chromatography

Both mixtures and solutions are composed of a least two substances that, although they are mixed together, do not combine chemically with each other. Sometimes there are many substances involved and in some cases these substances are very similar to each other.

In order to separate the components of a mixture or a solution it is necessary to find some difference in physical or chemical properties that the separation may be based on. Sometimes these differences are hard to find.

Chromatography is a means of separating the components of a multi-component system using a property difference that you would not imagine. The principle piece of apparatus is a piece of paper!

Paper Chromatography

2 250 ml erlenmeyer flasks
green grass
methyl alcohol
petri dish
filter paper

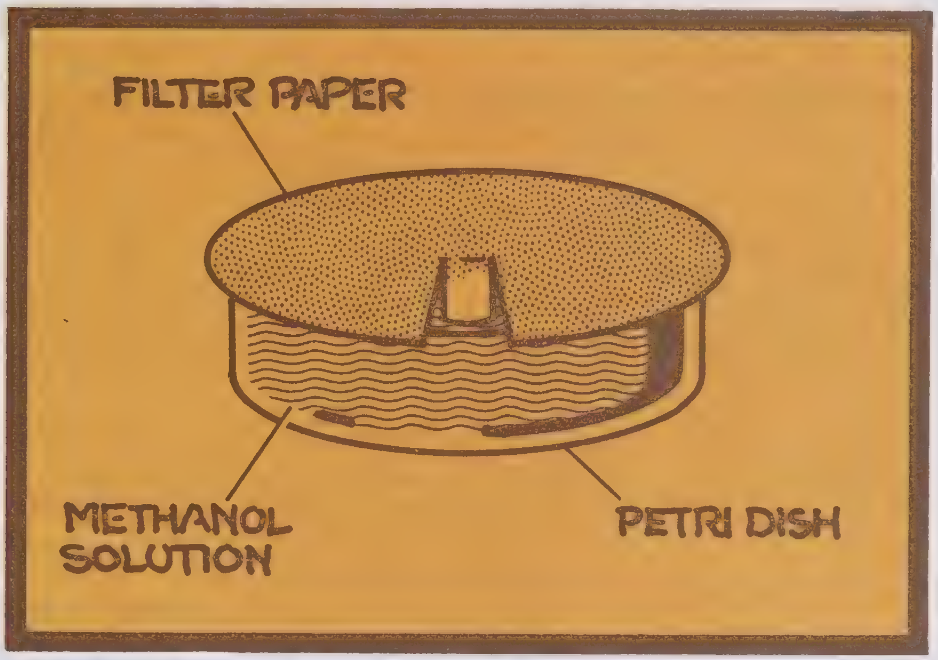
1. Put a handful of fresh green grass (or any other green plant material) into a 250 ml erlenmeyer flask. Add about 3 cm of methyl alcohol.

Warning

Do not bring methyl alcohol anywhere near an open flame.

2. Warm the flask of grass and alcohol on a hot plate until it begins to boil gently. Keep it boiling for about 3 min.

3. Pour off the liquid into a second 250 ml erlenmeyer flask. Continue to heat this liquid until the volume has been decreased to about half the original volume.
4. Pour about 20 ml of the liquid into a petri dish.
5. Cut a filter paper as shown in the diagram, fold down the tongue and place the paper over the petri dish so that the tongue dips into the liquid. Watch carefully.



Suggested Experiments

1. Try to extract the coloured material from a flower and separate it by chromatography.
2. Try using liquids other than methyl alcohol.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.

Water Pollution

Surely no introduction to the topic of water pollution is necessary. The problem is so great and so well published that little more can be added. It is really a question of survival.

Beat the problem or watch the planet die.

There is a tremendous amount that can be done here. One aspect of water pollution is simply finding out what the pollutants are. An example might be to determine the ammonia content of water. The amount of ammonia present in water can be related to the amount of organic wastes—frequently sewage—present in the water. This in turn can be related to disease organisms, bacteria and other factors.

Ammonia Content of Water

samples of water
pipette
test tubes
Nessler's reagent

1. Collect samples of water from around your locality. Rivers, lakes, streams, creeks — any natural water supply. If you like, select a stream and follow it through a town or city taking water samples every 100 m or so.
2. Back in the laboratory pipette 100 ml of each sample into separate test tubes.
3. Add 1 ml of Nessler's reagent to each test tube.
4. Arrange the test tubes in order of colour intensity. The most highly coloured one is the most polluted with ammonia. Try to relate the amount of pollution to the location that the sample came from.

Suggested Experiments

1. Expand the above experiment by mixing up a series of standard solutions of ammonia and Nessler's reagent to compare your samples with and hence actually measure the amount of ammonia present.
2. Test the samples for the presence of phosphates using the ammonium molybdate test.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Re-

prints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.

3. *Chemistry—Collected Experiments*. Nuffield Foundation. London: Penguin, 1967.

Air Pollution

Your generation will soon inherit a very imperfect world. Previous generations, particularly the last two or three, have created many serious problems that now threaten to overpower civilization. One of the most serious of these is air pollution.

Few of us have any real knowledge of air pollution. For most of us our knowledge has come from television, radio and the press. Just what is an air pollutant? Where do they come from? What sort of substances are they? Can they be dangerous? Can they be removed from the air? There is a long list of questions that needs to be examined.

Air Pollution

small aquarium
cardboard
flashlight
cotton balls
ammonium hydroxide
litmus paper

1. Obtain a small aquarium. Make a fairly tight fitting lid for it. Cover one end with a piece of heavy paper or cardboard that has a 0.5 cm hole in the centre of it. Mount a flashlight or similar light source about 60 cm from the aquarium so that the beam will hit the small hole and shine into the sample of confined gas. Cover the end of the aquarium where the beam exits with paper.
2. Soak a cotton ball in ammonium hydroxide, drop it in the aquarium and close the lid.
3. Darken the room and shine the beam of light through the aquarium. Observe carefully.
4. Gently open the lid and hold a piece of moist red litmus in the aquarium.
5. Flush out the aquarium and remove the cotton wad.
6. Blow some cigarette smoke into the box, close the lid, darken the room and shine the beam through the aquarium. Observe carefully.

7. Without disturbing the set up let it stand for an hour or more. Return to the apparatus and again examine the beam. Repeat this step several times.

Suggested Experiments

1. Make a *Cottrell precipitator* and show how it functions.
2. Collect samples of actual air-borne solids and try to identify their source.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, 1965.
3. Ontario Government Publications, Queen's Park, Toronto, Ontario.

Smoking

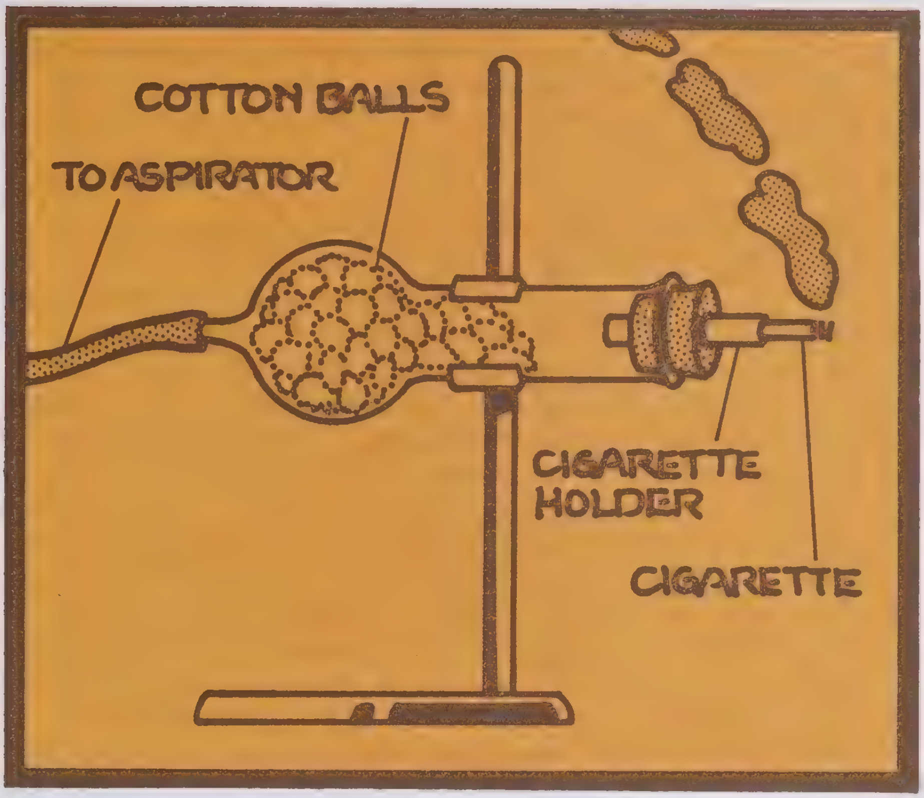
In recent years medical researchers have shown that cigarette smoking is hazardous to the health. A campaign to educate the nation about the dangers of this familiar habit has failed, largely because the majority of smokers do not want to listen. If you are a smoker this project may be for you.

Researchers have shown that cigarette smoking can increase the chances of your getting cancer. They have narrowed the guilty substances to the tars and nicotines that the smoker inhales. The more tar and nicotine — the better the odds for cancer.

Several factors influence the intake of tar and nicotine. This project deals with those factors.

Tar and Nicotine in Cigarettes

a smoking machine
9 mm tubing
rubber stopper
drying tube
aspirator
rubber tubing
cotton balls
filter cigarettes
non-filter cigarettes
ether



1. Make a "smoking machine" by inserting a 5 cm long section of 9 mm inside diameter tubing through a rubber stopper that will fit a drying tube; join the drying tube to an aspirator by means of a rubber tube. The finished equipment will look like the diagram.
2. Pack the drying tube with cotton balls and gently force a filter cigarette into the "mouth" of the machine. If the fit is too loose, wrap a short piece of transparent tape around the end of the cigarette to make it fit better. Turn on the water to start the suction and light the cigarette.
3. Adjust the suction so the cigarette is smoked at a fairly normal rate.
4. When the first cigarette is finished smoke a second and third cigarette.
5. After the third cigarette is done remove the cotton and wash it in small quantities of ether. Keep the ether in a test tube. Stopper the test tube.

Warning

Ether is highly inflammable and explosive. Do not have ether in an open container anywhere near a match or flame.

6. Now repeat the experiment using a non-filter cigarette. Make sure you use the same total amount of ether to wash the cotton when you are done.

Dyes and Dyeing

In about 1856 the first chemically produced dye was developed. The introduction of this dye opened the way to an entirely new industry, the manufacture of dyes.

Many of these chemically produced dyes originate from coal tar and as a result are referred to as *coal tar dyes*.

The coal tar itself does not contain dyes, but when coal tar is distilled fractionally naphthalene and anthracene are separated from the black sticky liquid. These are then used in the manufacture of what are known as *intermediates*. A series of complex reactions leads to a new group of compounds, each with a brilliant colour. These brilliant coloured compounds are then used as the dyes. Each has a complex formula and long chemical name which we shall not discuss here.

The dyes in use today are, of course, many and of a wide variety of colours. They are, however, grouped into several different classes depending upon how they are used and the type of material they are used on.

Direct dyes are used on all natural fibres like wool, cotton or silk. Acid dyes give bright fast colours to wool and silk; however, they do not cling to cotton or linen. Mordant dyes are those which require some sort of chemical treatment before the material is actually placed in the dye bath.

An entirely different group of dyes, widely used before the development of chemical dyes, were those extracted from leaves, roots, berries and bark. Some were extracted from animal products and a few minerals were used.

Because of the availability of leaves, roots or bark we shall look more closely at these dyes instead of the more complex chemical dyes. In the experiment outlined below you will attempt to make and use several dyes, each made from a readily available plant or plant substance.

Preparation and Use of an Onion Dye

onion skins
alum
cream of tartar

1. Boil the brown peel of ripe onions in water for 30 min; then strain off the coloured liquid for the dye.
2. Prepare a mordant (dye fixer) of alum. This is done by dissolving completely about 2.2 g of alum and 0.75 g of cream of tartar in 1 ℓ of distilled water.
3. Wet the material to be dyed (wool or cotton) in the mordant. This is done by immersing the material in the mordant bath and stirring it gently. Gradually raise the temperature but not to boiling point. Simmer for about an hour.
4. Remove the material from the mordant bath and place it (while wet) in the dye bath. Continue simmering until the desired colour (a yellow shade) is imparted to the material.
5. Remove from the dye bath, rinse and allow the cloth to dry. Wash the cloth now to see if the colour will remain in it.

Suggested Dyes

1. Spruce dye
2. Birch dye
3. Oak bark dye
4. Tea dye
5. Butternut dye
6. Goldenrod dye

Sources of Further Information

1. "Dye Plants and Dyeing," *Plants and Gardens*. Brooklyn Botanic Garden, 1000 Washington Ave., Brooklyn, New York.
2. *Modern Chemistry*. Metcalfe, H.C. et al. New York: Holt, Rinehart and Winston, 1970.
3. *Dyes from Plants*. Seonaid Robertson. New York: Van Nostrand Reinhold, 1973.

Electroplating

You have surely at some time admired the nice shiny finish on water taps, the bumpers of cars or even the top of your ballpoint pen. That finish was certainly never painted on. Have you ever thought how great it would be if you could produce a finish like that on the blade of your pocket knife, your car key or your home-made ring?

The nice shiny finishes mentioned above are all produced by a process called electroplating.

In electroplating a layer of metal is deposited by electrolysis. The object being plated is the cathode. The electrolytic tank or bath in which the plating is carried out contains a solution of a salt of the metal to be deposited. For example, if the metal to be deposited is copper, the bath must be a solution of copper salt.

How to Electroplate

dilute HCl
steel wool
ammeter
rheostat
$CuSO_4$
ethyl alcohol
concentrated H_2SO_4
gelatin

The success of an electroplating depends a great deal upon several things:

1. The object being plated must be thoroughly cleaned. This is accomplished by first washing the object being plated in a dilute hydrochloric acid solution then rinsing with distilled water followed by a rubbing with a fine steel wool or jeweller's rouge. This should be repeated several times and the object finally rinsed with distilled water.

Warning

The object being plated must not be touched after cleaning.

2. The current should not exceed 0.5 A. An ammeter and rheostat are placed in series in the circuit to control the current.
3. The bath solution must be carefully prepared. Follow this recipe:

   70 g of $CuSO_4$ in 500 ml of distilled water
   24 ml of ethyl alcohol

   Dissolve each of the above in the water then add 14 ml of concentrated H_2SO_4. Finally, add a few grains of clear gelatin to produce a smooth finish.
4. Set up the apparatus as illustrated in the diagram and allow the current to flow until a reasonable layer of copper has been deposited.

Suggested Experiments

Plate with other metals such as chromium, silver or tin.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
2. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
3. *Modern Chemistry*. Metcalfe, H.C. et al. New York: Holt, Rinehart and Winston, 1970.

Electrolysis

Electrolysis is the name given to decomposition reactions in which electric current is passed through a substance in either a molten or dissolved state. The process works only on those substances that produce *ions* when molten or dissolved. These charged particles migrate through the solution or fused liquid under the influence of the electric field. When an ion reaches the oppositely charged electrode its charge is nullified and some new substance is liberated at the electrode. Sometimes the product of the electrolysis reacts with either the solvent or the electrode. Sometimes the product reacts with both.

Electrolysis is normally carried out with low voltage direct current.

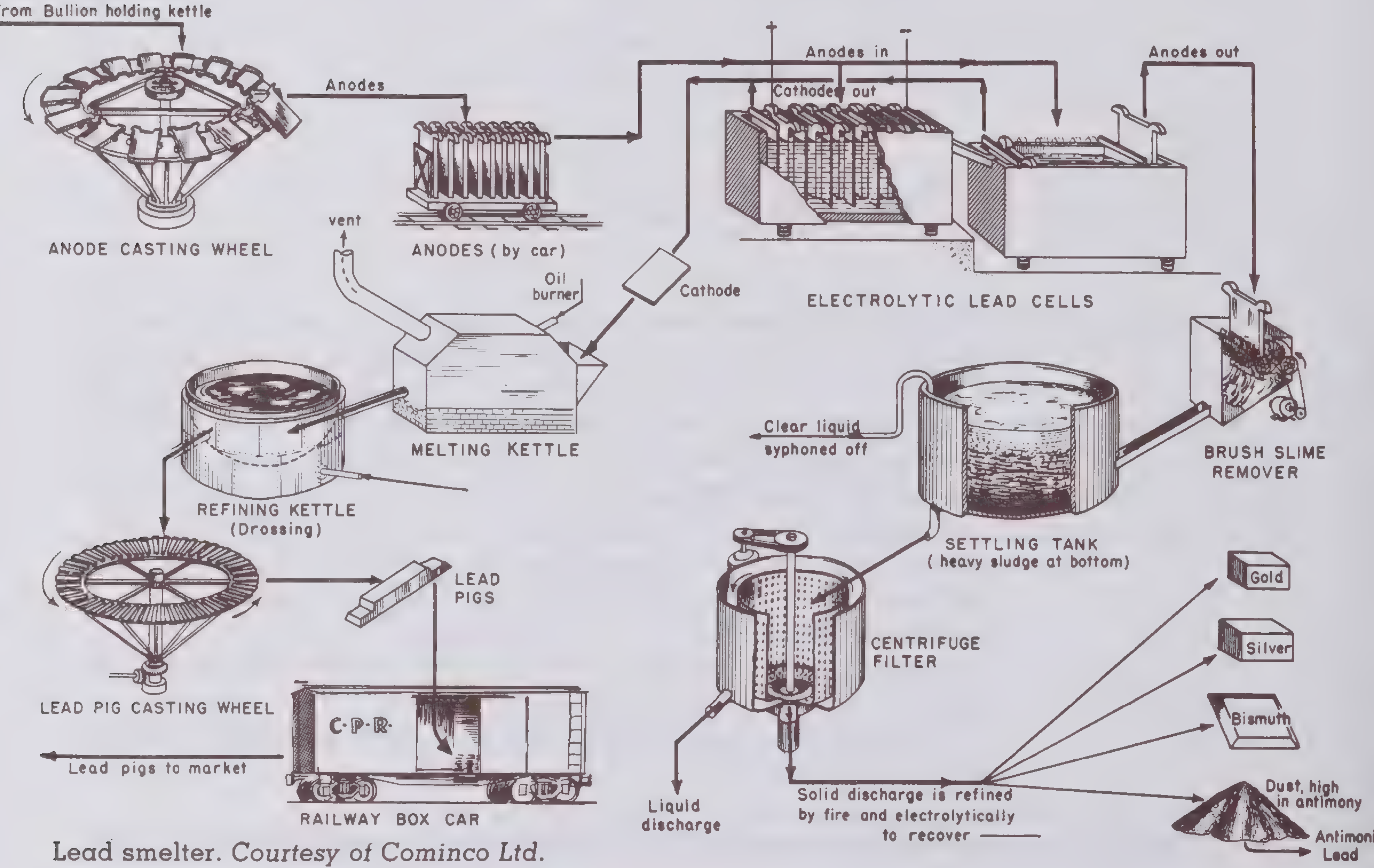
Chlorine gas is an example of a substance produced commercially by electrolysis.

Electrolysis of Iodide

U-tube
0.5 M potassium iodide
eye dropper
phenolphthalein
starch

1. Fill a U-tube three-quarters full with a 0.5 M solution of potassium iodide.

2. Lower two carbon electrodes into the solution, one in each arm. Connect them to a variable (DC) power source.
3. Turn on the power supply and let the electrolysis proceed for 5 to 10 min at 8 to 10 V. Note and record any products and any colour changes that occur.
4. Using an eye dropper remove 2 to 3 ml of solution from each side of the U-tube.
5. Add a few drops of phenolphthalein solution to the colourless sample.
6. Add a few drops of starch solution to the coloured sample.


rom Bullion holding kettle
Anodes
ANODE CASTING WHEEL
ANODES (by car)
vent
Oil burner
Cathode
Anodes in
Cathodes out
Anodes out
ELECTROLYTIC LEAD CELLS
MELTING KETTLE
REFINING KETTLE (Drossing)
Clear liquid syphoned off
BRUSH SLIME REMOVER
SETTLING TANK (heavy sludge at bottom)
LEAD PIGS
LEAD PIG CASTING WHEEL
C·P·R·
Lead pigs to market
RAILWAY BOX CAR
CENTRIFUGE FILTER
Gold
Silver
Bismuth
Dust, high in antimony
Antimoni
Lead
Liquid discharge
Solid discharge is refined by fire and electrolytically to recover

Lead smelter. *Courtesy of Cominco Ltd.*

Suggested Experiments

Replace the potassium iodide solution with copper chloride solution. You will have to be more cautious this time and you will have to change Steps 4, 5 and 6 to compensate for this different solution.

Sources of Further Information

1. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.

2. *Chemistry Man's Servant*. Fludner amd Tiechman. Allyn Bacon.
3. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
4. *Chemical Process Industries*. Shreve. New York: McGraw-Hill, 1967.
5. *Chemistry – Collected Experiments*. Nuffield Foundation. London: Penguin, 1967.

Other Projects

The preceeding 21 projects are not a complete list. There are obviously many other things you can do. If you wish to strike out on your own, then consult with your teacher. Perhaps she/he can help you select a topic. Perhaps a second list of projects will help you find something you can do.

1. Manufacture glass from basic materials. Colour the glass with other materials.
2. Investigate the makeup and effectiveness of household cleaners.
3. Manufacture soap from basic materials. Colour and perfume the product.
4. Examine foods for acidity and try to isolate or identify the acids present.
5. Separate caffein from tea or coffee.
6. Investigate the chemistry used by police in identifying blood smears or different soil types.
7. Investigate metal alloys by preparing some common ones such as solder or Wood's metal.
8. Grow a silicate garden and investigate the nature of silicates and silicones.
9. Investigate batteries and determine how a chemical reaction can produce an electric current. Look into fuel cells.
10. Prepare textile fibres from basic materials. Learn to identify textile fibres using chemical tests.
11. Identify the test for Vitamin C. Test various fruit juices and drinks for the amount of Vitamin C present.
12. Investigate the chemistry of building materials such as cement and plastic.
13. Make paint, enamels or lacquers from basic ingredients.
14. Investigate detergents and how they work. How do soaps differ from detergents?
15. Identify the chemicals needed by a growing plant.

Perhaps the sort of projects given here still don't turn you on. Perhaps what you need is not a project title, but a question that needs an answer. Here are some questions that relate to chemistry or might involve some chemistry to answer them.

1. Why does a candle go out when confined?
2. How can one determine what substances are present in a rock?
3. Do substances with the same colour contain the same chemicals?
4. Is a given element the same colour in all compounds that it forms?
5. How much carbon dioxide is present in the gas exhaled from the lungs? Does this amount change with various stages of activity?
6. How can one remove the perfume from flowers?
7. Can fruit and vegetable colouring be used for dyes or indicators?
8. What substances are dissolved in our drinking water?
9. What substances are present in crude oil?
10. What substances are present in coal?
11. How strong are soap bubbles? How can they be made stronger?
12. Are bleaches and dyes hard on hair? Do these substances damage hair if used repeatedly?
13. Why do flashlight batteries go dead? Can they be brought back to life again?
14. What is a glue or adhesive? How do they work?
15. Can any substance be both an acid and a base?

Sources

The library is filled with information but obviously some sources are better than others for chemistry. Here are five of the best sources for chemistry projects.

1. *Nuffield Science Series*. Toronto: Longmans Canada, 1967.
2. *Sourcebook for the Physical Sciences*. Joseph, A. et al. New York: Harcourt, Brace and World, 1961.
3. *Tested Demonstrations*. Alyea, H.N., and Dulton, F.B. ACS. Reprints from *The Journal of Chemical Education*, American Chemical Society, Easton, 1965.
4. *Griffin Science Series*. Middlesex, England: Griffin and George, 1971.
5. *The Chemical Formulary*. Bennett, H. New York: Chemical Publishing Co., 1945.

Element	Symbol	Atomic Number	Element	Symbol	Atomic Number
Actinium	Ac	89	Mercury	Hg	80
Aluminum	Al	13	Molybdenum	Mo	42
Americium	Am	95	Neodymium	Nd	60
Antimony	Sb	51	Neon	Ne	10
Argon	Ar	18	Neptunium	Np	93
Arsenic	As	33	Nickel	Ni	28
Astatine	At	85	Niobium	Nb	41
Barium	Ba	56	Nitrogen	N	7
Berkelium	Bk	97	Nobelium	No	102
Beryllium	Be	4	Osmium	Os	76
Bismuth	Bi	83	Oxygen	O	8
Boron	B	5	Palladium	Pd	46
Bromine	Br	35	Phosphorus	P	15
Cadmium	Cd	48	Platinum	Pt	78
Calcium	Ca	20	Plutonium	Pu	94
Californium	Cf	98	Polonium	Po	84
Carbon	C	6	Potassium	K	19
Cerium	Ce	58	Praseodymium	Pr	59
Cesium	Cs	55	Promethium	Pm	61
Chlorine	Cl	17	Protactinium	Pa	91
Chromium	Cr	24	Radium	Ra	88
Cobalt	Co	27	Radon	Rn	86
Copper	Cu	29	Rhenium	Re	75
Curium	Cm	96	Rhodium	Rh	45
Dysprosium	Dy	66	Rubidium	Rb	37
Einsteinium	Es	99	Ruthenium	Ru	44
Erbium	Er	68	Samarium	Sm	62
Europium	Eu	63	Scandium	Sc	21
Fermium	Fm	100	Selenium	Se	34
Fluorine	F	9	Silicon	Si	14
Francium	Fr	87	Silver	Ag	47
Gadolinium	Gd	64	Sodium	Na	11
Gallium	Ga	31	Strontium	Sr	38
Germanium	Ge	32	Sulfur	S	16
Gold	Au	79	Tantalum	Ta	73
Hafnium	Hf	72	Technetium	Tc	43
Helium	He	2	Tellurium	Te	52
Holmium	Ho	67	Terbium	Tb	65
Hydrogen	H	1	Thallium	Tl	81
Indium	In	49	Thorium	Th	90
Iodine	I	53	Thulium	Tm	69
Iridium	Ir	77	Tin	Sn	50
Iron	Fe	26	Titanium	Ti	22
Krypton	Kr	36	Tungsten	W	74
Lanthanum	La	57	Uranium	U	92
Lawrencium	Lw	103	Vanadium	V	23
Lead	Pb	82	Xenon	Xe	54
Lithium	Li	3	Ytterbium	Yb	70
Lutecium	Lu	71	Yttrium	Y	39
Magnesium	Mg	12	Zinc	Zn	30
Manganese	Mn	25	Zirconium	Zr	40
Mendelevium	Md	101			

PERIODIC TABLE OF THE ELEMENTS

ATOMIC NUMBER 7
SYMBOL **N**
ATOMIC WEIGHT 14.0
VALENCE (OXIDATION) NUMBERS 3 5

IA	IIA	III B	IV B	V B	VI B	VII B	VIII B	VIII B	VIII B	I B	II B	III A	IV A	V A	VI A	VII A	Inert Gases
1 1 **H** 1.0																1 1 **H** 1.0	2 **He** 4.0
3 1 **Li** 6.9	4 2 **Be** 9.0											5 3 **B** 10.8	6 4 2 **C** 12.0	7 3 5 **N** 14.0	8 2 **O** 16.0	9 1 **F** 19.0	10 **Ne** 20.2
11 1 **Na** 23.0	12 2 **Mg** 24.3											13 3 **Al** 27.0	14 4 **Si** 28.1	15 5 3 **P** 31.0	16 6 4 2 **S** 32.1	17 1 5 7 **Cl** 35.5	18 **Ar** 39.9
19 1 **K** 39.1	20 2 **Ca** 40.1	21 3 **Sc** 45.1	22 3 4 **Ti** 47.9	23 5 4 3 2 **V** 51.0	24 3 6 2 **Cr** 52.0	25 2 4 7 **Mn** 54.9	26 3 2 **Fe** 55.9	27 2 3 **Co** 58.9	28 2 3 **Ni** 58.7	29 2 1 **Cu** 63.5	30 2 **Zn** 65.4	31 3 **Ga** 69.7	32 4 **Ge** 72.6	33 3 5 **As** 74.9	34 4 **Se** 79.0	35 1 5 **Br** 79.9	36 **Kr** 83.7
37 1 **Rb** 85.5	38 2 **Sr** 87.6	39 3 **Y** 89.9	40 4 **Zr** 91.2	41 5 3 **Nb** 92.9	42 6 3 **Mo** 96.0	43 7 **Tc** 99	44 3 4 8 **Ru** 101.7	45 3 4 **Rh** 102.9	46 2 4 **Pd** 106.7	47 1 **Ag** 107.9	48 2 **Cd** 112.4	49 3 **In** 114.8	50 4 2 **Sn** 118.7	51 3 5 **Sb** 121.8	52 4 6 2 **Te** 127.6	53 1 5 **I** 126.9	54 **Xe** 131.3
55 1 **Cs** 132.9	56 2 **Ba** 137.4	57 3 **La** 138.9	72 4 **Hf** 178.6	73 5 **Ta** 180.9	74 6 **W** 183.9	75 7 4 **Re** 186.3	76 4 6 8 **Os** 190.2	77 3 4 **Ir** 193.1	78 4 2 **Pt** 195.2	79 3 **Au** 197.2	80 2 1 **Hg** 200.6	81 1 3 **Tl** 204.4	82 2 4 **Pb** 207.2	83 3 5 **Bi** 209.0	84 2 4 **Po** 210	85 1 **At** 210	86 **Rn** 222
87 1 **Fr** 223	88 2 **Ra** 226.1	89 3 **Ac** 227															

58 3 4 **Ce** 140.1	59 3 **Pr** 140.9	60 3 **Nd** 144.3	61 3 **Pm** 147	62 3 2 **Sm** 150.4	63 3 2 **Eu** 152.0	64 3 **Gd** 156.9	65 3 4 **Tb** 159.2	66 3 4 **Dy** 162.5	67 3 **Ho** 164.9	68 3 **Er** 167.2	69 3 **Tm** 169.4	70 3 2 **Yb** 173.0	71 3 **Lu** 175.0
90 4 **Th** 232.1	91 5 4 **Pa** 231	92 6 4 3 5 **U** 238.0	93 5 3 4 6 **Np** 237	94 4 3 5 6 **Pu** 244	95 3 4 5 6 **Am** 243	96 3 **Cm** 247	97 3 4 **Bk** 247	98 3 **Cf** 252	99 3 **Es** 254	100 3 **Fm** 257	101 3 **Md** 257	102 3 **No** 255	103 3 **Lr** 256

Answers

Page 48
1. 1 300 00 J
3. 3340 J
4. 17.8 g
5. 3.3 g
6. 46 743 J
7. 199.5 g

Page 68
11. 37.5 g/100g
12. 44 g/100 g
13. 9 g/100g
15. 19.3 g
16. 5 g
19. (a) 45 g/100 g; 31.5 g/100 g
    (b) 53°C; 33°C
    (d) 47.5 g
    (e) 10.4 g
    (f) 64.5 g

Page 102
1. 3.09 ℓ
2. 5.34 ℓ
3. 8.67 ml

Page 105
1. 8.64 ℓ
2. 0.57 ℓ
3. 1736 ℓ
4. 824 ml; 12.23 ℓ; 267 ml; 5.09 x 10 ml; 0.56 ℓ
5. (a) 245.7 ml
   (b) 0.60 g
   (c) 42.7 ml
6. 0.343 g

Page 127
1. (a) 23 g
   (b) 20 g
   (c) 33.75 g
   (d) 1.2 g
2. (a) 2 mol
   (b) 1.25 mol
   (c) 0.25 mol
   (d) 4.9 mol
   (e) 10 mol

Page 127
3. (a) 36.5 g
   (b) 98 g
   (c) 98 g
   (d) 342 g
   (e) 60 g
   (f) 46 g
4. (a) 36.5 g
   (b) 24.5 g
   (c) 245 g
   (d) 6.9 g
5. (a) 0.5 mol
   (b) 0.5 mol
   (c) 1.78 mol
   (d) 0.2 mol
   (e) 32.9 mol

Page 131
1. (b) 2.17 g
   (c) 0.018 mol; 0.027 mol
   (d) 0.009 mol; 0.013 mol
7. 22.3 ℓ
8. 2.9 x 10^{23}
9. (a) 12.04 x 10^{23}
   (b) 48.16 x 10^{23}
   (c) 48.16 x 10^{23}
10. (a) 2.5 mol
    (b) 9.04 mol
    (c) 0.5 mol
11. (a) 136 g
    (b) 87 g
    (c) 62 g
12. 214.2 mm; 428.4 mm; 107.1 mm
13. 32 g
14. 27.3%
15. 30%
16. 25.5%
17. 0.99
18. 220
19. 28 g

Page 232
1. (a) 58.5 g/1 ℓ
   (b) 1275 g/3 ℓ
   (c) 196.3 g/1.5 ℓ
   (d) 7.4 g/0.25 ℓ
   (e) 0.034 g/0.1 ℓ
2. 83.3 ml/1 ℓ
3. 55.5 ml/1 ℓ
4. 147 ml/0.5 ℓ
5. 625 ml/2 ℓ
6. 67 ml/1 ℓ

Page 238
1. 0.75 M
2. 0.75 M
3. 0.166 M
4. 2 x 10^{-2}
5. 37 g

Page 248
5. 4 mol
6. 0.51 g
7. 3.42 ℓ

Page 250
6. 0.25 g
7. 5.6 ℓ

Page 256
1. 700 ℓ
2. 3.03 ℓ
3. 6.85 ℓ
4. 22.8 ℓ
5. 5 x 10^6 ℓ
6. 9 g:27 ml
7. 22.4 ℓ

Page 262
1. 68 g
2. 3.9 g; 2.7 ℓ; 0.12 mol
3. 6.8 g
4. 186.6 g
5. \$6.30; \$2.45; \$1.50

Page 267
1. 5 mol
2. 3 mol
3. 6.6 mol
4. 9 mol
5. 2 mol

6. 2.5 mol
7. 160 g
8. 40.8 g
9. 2 g; 2 g; 2 g
10. 0.4 g
11. 11.25 g
12. 92 g
13. 131 g
14. 18 250 g
15. 700 kg
16. 1.45 g
17. 1607 g
18. 1.42 g
19. 1244 ℓ
20. 9333 ℓ
21. 2.5 ℓ
22. (a) 200 ℓ
    (b) 10 ℓ
23. 500 ℓ; 1500 ℓ
24. 500 ℓ
25. 25 ℓ

Index

acetylene, 323
acids, 219, 223, 238, 243
acid strength, 226, 419
acid, ternary, 224
alcohol, 330, 333
aldehyde, 334, 335
alkane, 318, 323
alkene, 320, 324, 327
allotrope, 314
alkyl group, 324
alkyne, 321, 323, 327
alpha particle, 151
amalgam, 61
amorphous, 314
Arrhenius, Svante, 179
atom, 119, 122
atomic diagrams, 173
atomic dimensions, 159, 290
atomic mass, 125, 160, 168, 197, 290
atomic model, 144, 157, 162, 169
atomic number, 166
Atomic Numbers, Law of, 166, 290
atomic structure, 168
Avogadro, Amedeo, 117
Avogadro's Hypothesis, 117

Baekeland, Leo, 348
balances, 384
barometer, 90
Bartlett, Neil, 305
bases, 221
Berzelius, J., 122
Becquerel, H., 151, 197
binary, 184, 186
black box, 138
bleaches, 364
Bohr atom, 160
Bohr, Neils, 160
bonding, 191, 415
Boyle's Law, 96, 98, 130
Boyle, Robert, 96
bromide, test for, 288
bromine, 278
bunsen burner, 382
burette, 236

calorimetry, 42
carbon, 310, 315
carbon dioxide, 252, 254
carboxylic acid, 236
catalysis, 212
cathode rays, 146
chains, carbon, 317
charge, 141
Charles' Law, 130
chemical change, 40
chemistry, 85
chloride, test for, 288
chlorine, 277, 362
chromatography, 424
colloid, 58
colloidal state, 59
conclusions, 23
condensation polymer, 344
conditions that matter, 31
conductivity, 176, 226
conservation of mass, 84, 130, 186
Constant Composition, Law of, 84, 130
cosmetics, 412
coulomb, 158
Coulomb's Law, 156
covalent bond, 192, 416
covalent radius, 297
Crookes, William, 147
crystals, 422
Curie, M. and P., 151, 197

Dalton, John, 130
Definite Proportion, Law of, 84, 130
density, 300
density of air, 89, 160
derivatives, organic, 339
destructive distillation, 311, 313
diamond, 314
diffusion of gases, 109
displacement reactions, 242
dissolve, 49
distillation, 57, 316
drugs, 403
dry ice, 263
dyes, 430

electrolysis, 178, 189, 433
electron, 149, 157
electron configuration, 170, 174, 176
electronegative, 289
electroplating, 431
electropositive, 289
Electrostatics, Law of, 141
element, 122
end point, 229
endothermic, 42
energy levels, 172
equations, 136, 189, 244, 254
esters, 431
eudiometer, 112
exothermic, 42

fertilizers, 376
fibres, 378
filtration, 387
flame tests, 371
foods, 405
force (newton), 92
formula, 184
functional groups, 329, 342, 350

gas model, 110
gas pressure, 90, 92
graphical analysis, 295
graphite, 314
group, 290

Hahn, Otto, 199
halogen, 275, 292
heat quantity, 43
heterogeneous, 52
homogeneous, 52
hydrides, 292
hydrocarbons, 317
hydrocarbon oxidation, 341
hydrogen, 246, 252
hydrogen halides, 284
hydronium ion, 227
hypothesis, 34

Ideal Gas Law, 128
indicators, 419
interpretation, 32
iodide, 288
iodine, 280
inert gas, 175, 291, 305
ionization energy, 301
ions, 179
isomers, 335
isotopes, 195, 201

joule, 43

Kelvin temperature, 101
ketone, 334
kinetic energy, 108

laxative, 377
Lewis diagram, 193

manometer, 118
mechanical mixture, 52, 53
Meitner, Lise, 199
melting temperature, 296, 398
Mendeleev, Dmitri, 272, 305
metals, 272, 305, 368
methyl group, 324
Meyer, Earl, 32
Meyer, Lothar, 305
milk, 407
molar volume, 127, 130
molarity, 229
mole, 123
molecular dimensions, 413
molecular mass, 125, 409, 410
molecular model, 137
molecule, 119
monomer, 344
Moseley, H., 166, 301

neutralization, 227
neutron, 168
Newlands, J., 305
Newton, 94
noble gas, 305
non-metal, 305
nucleus, 155

observation, 20
organic chemistry, 310, 350
organic compounds, 342
oxidation, 351
oxidative addition, 324
oxygen, 213

paraffin, 318, 324
partial pressure, 117, 138
particle model, 106
pascal, 92
Periodic Law, 301
periodic table, 301, 303
peroxides, 211
petroleum, 317, 421
phase, 50
Phlogiston Theory, 21, 72
photography, 365
physical change, 37, 39
plastic, 343
polymer, 344
precision, 83
Priestley, J., 208, 264
pressure, 92
properties, 57
proton, 150
pulp and paper, 417

qualitative analysis, 371, 322, 375

radioactive, 152
radioactivity, 197
reduction, 251
regularity, 28
Röentgen, Wilhelm, 198
Rutherford, Ernest, 152
Rutherford atom, 152, 199

salt, 228
sclerometer, 368
shells, 172
soap, 360
solubility, 62
solubility chart, 67
solubility curve, 63, 65, 66
solute, 49
solution, 49, 61, 62
solvent, 49
specific heat, 43
spectator ions, 189
stability, 182
Stahl, G., 21
states of matter, 39
static electricity, 140
stoichiometry, 258
STP, 103, 105
sub-atomic particle, 149
symbol, 123

temperature, 108
tensile strength, 369
titration, 229, 233, 239
theory, 117, 130
Thomson, J. J., 144
Thomson atom, 144

valence, 182, 299
valence shell, 298
Van Guericke, Otto, 90
vapour pressure, 120
volume measurement, 389

water, hardness of, 356

x-rays, 151

1 2 3 4 5 6 7 8 9 0

DATE DUE SLIP

EDUC AP 6 '83

RETURN APR 6 '83

DUE EDUC AUG 19 '83

AUG 19 RETURN

DUE EDUC NOV 24 '83

DUE EDUC DEC 1 '83

DUE EDUC DEC 7 '83

DUE EDUC DEC 15 '83

DEC 15 RETURN

DUE EDUC MAR 30 '86

APR 04 RETURN

DUE EDUC NOV 28 '86

DEC 05 RETURN

F. 255

QD 33 C86 1976
COURNEYA DAVE
THE NATURE OF MATTER/

39300189 CURR

LIBRARY USE ONLY

000004769121

QD 33 C86 1976
Courneya, Dave.
The nature of matter /

39300189 CURR

RECOMMENDED FOR USE
IN ALBERTA SCHOOLS

B21870

ISBN 0-669-00542-8